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STUDIES OF RATES OF CONVERSION AND POPULATIONS OF VARIOUS CONFORMATIONS OF SATURATED RING COMPOUNDS BY N.M.R.

I. CHLOROCYCLOHEXANE AND BROMOCYCLOHEXANE[1]

L. W. REEVES AND K. O. STRØMME[2]

ABSTRACT

The populations of equatorial and axial halogen forms of chloro- and bromo-cyclohexane have been studied by proton magnetic resonance between 25° and −104° C in carbon disulphide solution. From the equilibrium constants obtained the free energy change between the two conformations is obtained. The bromo compound has 82% of the equatorial halogen form and the chloro compound 77%. Estimates of the populations are also available for room temperature from the chemical shift measurements. The energy barrier for conversion between the two forms is approximately 10,850 calories. A simple Lennard-Jones potential with the available estimates of van der Waals radii is used to calculate the differences in energy of the two forms.

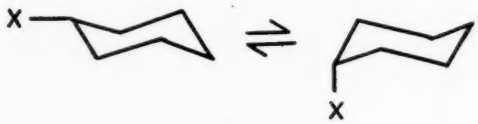
The difference in nuclear shielding between the axial and equatorial protons in acetylated sugars was first detected by Lemieux, Kullnig, Bernstein, and Schneider (1). This has become a very useful method of determining the conformation of sugars and has been followed up by studies of trioxanes (2) and the isomeric hexachlorocyclohexanes (3). In almost all these compounds the conformations are locked by large groups so that in most cases distinct isomers can be isolated. In the case of the hexachlorocyclohexanes the theoretically possible isomers are reduced to just a few known compounds by the fact that steric hindrance renders almost all of these conformations unfavorable (4).

In contrast to the highly substituted derivatives, the simpler cyclohexanes do not show proton resonance spectra with distinct signals for axial and equatorial protons (5). The lack of large substituted groups allows the rapid interconversion of the various conformations and only the averaged shielding parameter is effective (6). The proton resonance technique has proved a useful method of measuring the lifetime of nuclei in distinct chemical environments (7, 8, 9, 10) and the present and future studies of these systems will be aimed at measuring the rates of interconversion between conformations and the population of the conformers in equilibrium at temperatures where distinct proton resonances can be obtained for the individual species.

Such studies will give enthalpy differences and energy barriers for the internal conversion over series of related compounds.

[1]*Manuscript received April 19, 1960.*
Contribution from the Department of Chemistry, University of British Columbia, Vancouver 8, British Columbia.
[2]*Permanent address: Chemistry Department, University of Oslo, Blindern, Norway.*

In this present paper two compounds of the type



are studied. The equilibrium populations are obtained in one solvent at various temperatures. An estimated rate is obtained at one temperature but the study of variation of solvent will be left for later work. The influence of solvent on such isomeric equilibria in solution has been dealt with by Bernstein and Powling (11) and Reeves (12, 10).

EXPERIMENTAL

Analytical grade chlorocyclohexane and bromocyclohexane were used without further purification. The solutions were prepared with redistilled carbon disulphide, which is the most suitable solvent for low-temperature work and has the advantage of no external dipole moment and no proton resonance. It is intended to vary polarity of solvent in later work.

The solutions prepared had the following composition:

(A) mole% $C_6H_{11}Cl$ = 54.5, mole% CS_2 = 37.4, mole% tetramethyl silane = 8.1;

(B) mole% $C_6H_{11}Br$ = 55.4, mole% CS_2 = 41.2, mole% tetramethyl silane = 3.4.

A 40 Mc/sec Varian V4300 high-resolution spectrometer was used for the work. The low-temperature arrangement used was a modification of an apparatus described by Schneider, Bernstein, and Pople (13). The arrangement will be described more completely later. The samples are contained in 3-mm I.D. tubes which are sealed off in vacuum and have two bearings: the upper one, a tungsten wire, sealed into glass, and the lower one, a glass tit bearing, which lowers into a drilled teflon plate. The sealing and accurate alignment of bearings is achieved by use of a microglass lathe to be described by one of us (L. W. R.).

The tube is contained in a Dewar arrangement previously described (13). A 15-mm Varian insert is used and there is Dewar insulation between the insert coil and the cooled sample. The small air space between the Dewar and the insert wall is circulated with dry nitrogen at room temperature so that no cooling or condensation can occur even on the outside of the transparent Dewar. Spurious water peaks are thus avoided.

Cooling is achieved by a heater placed in a 25-liter metal Dewar of liquid nitrogen with exit tube and blowoff valve. In cases where it is necessary to heat the sample at any time, a small heating coil in the nitrogen stream is used. The cooled nitrogen stream spins the sample from tangential holes striking the shaped glass rod top of the sample tube. Since a more or less constant spinning rate is desirable the faster streams of nitrogen obtained at low temperatures are bled off in a side tube and only part of the stream is used for spinning and cooling.

A previous study of methyl nitrite at low temperatures demonstrates the reproducibility of the apparatus (12(*b*)). Providing the viscosity of materials is favorable, high-resolution spectra comparable to those taken at room temperature can be obtained down to $-180°$ C and in this work spectra were measured in the region as low as $-120°$ C. The temperature control as indicated by a copper constantan thermocouple in the exit stream near the sample is approximately $\pm 1°$ C down to $-80°$ C and $\pm 1.5°$ C below this, over 15-minute periods.

RESULTS

Spectra of solution A and solution B at 28° C are shown in Fig. 1. Spectrum (*a*) corresponds to bromocyclohexane in CS_2 and spectrum (*b*) corresponds to chlorocyclohexane in CS_2. Peak II is the proton resonance in the —CHX— proton and peak I in the $—C_5H_{10}—$ protons. The internal standard tetramethyl silane is shown to high field as peak 0. The spectrum is calibrated by 50, 100, 150, and 200 c.p.s. side bands in some

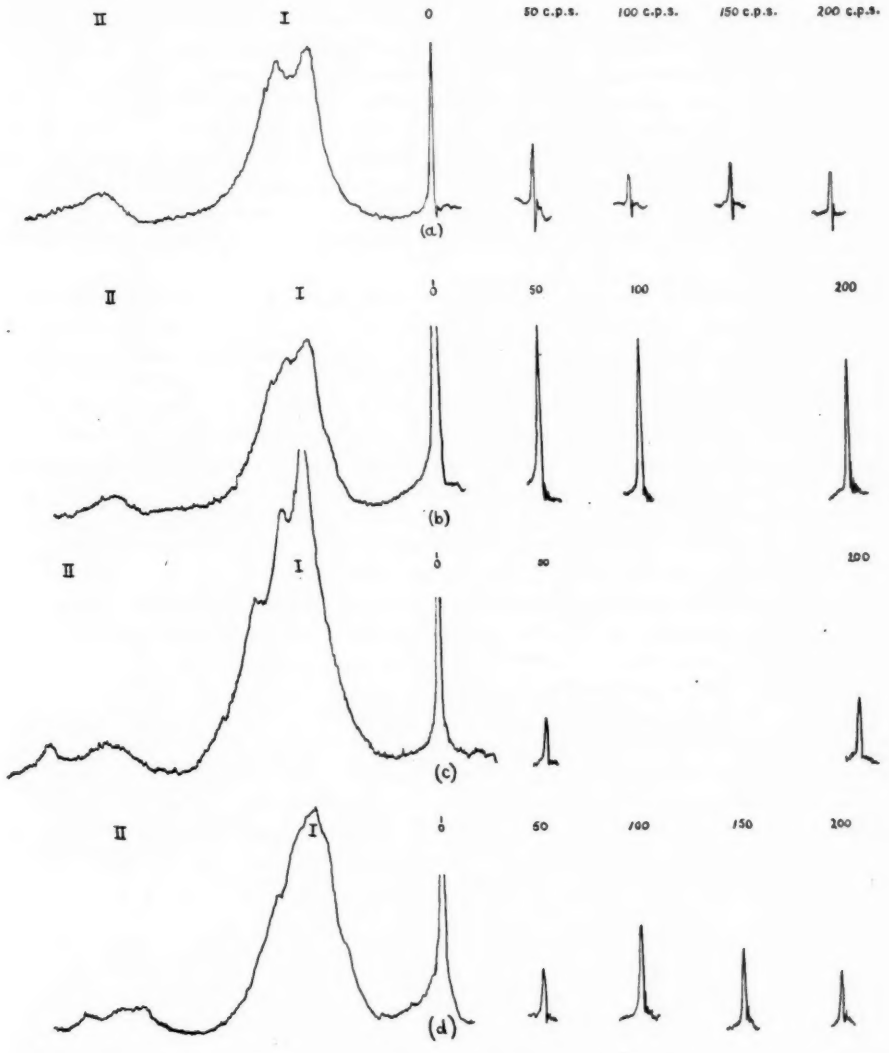


FIG. 1. (*a*) Spectrum of bromocyclohexane at 40 Mc/sec in CS_2 with $Si(Me)_4$ standard signal. Solution B 28° C. (*b*) Spectrum of chlorocyclohexane solution in CS_2 at 28° C. (*c*) Spectrum of bromocyclohexane at −104.5° C. (*d*) Spectrum of chlorocyclohexane at −91° C.

cases, as shown in Fig. 1. The irregular base line which occurs during the seep between side bands due to the process of switching off the audio oscillator and resetting at a new value is caused by side bands of the $—C_5H_{10}—$ peaks. The base line has therefore been removed in most cases. The lack of ringing off the first tetramethyl silane peak is due to a prompt addition of 50 c.p.s. side bands after passing through this signal. The instrumental resolution is better seen from the audio side band peaks.

The —CHX— signal is seen to be a broad band approximately 12 c.p.s. in half width. This arises from a number of similar coupling constants with nuclei which are considerably chemically shifted ($—C_5H_{10}—$) and is characteristic of similar situations which arise in substituted hydrocarbon chains (3, p. 235–238). The fine structure of this peak can to a certain extent be resolved with higher gains and slower passage through resonance but this has not been relevant to the present study. The signal I in spectrum (*a*) for bromocyclohexane consists of two main broad peaks. The low-field peak is due to the $—CH_2—$ groups on each side of the —CHX— grouping and the high-field part of this group of lines is due to the remote $—(CH_2)_3—$ protons from the —CHX— grouping.

The main features of the chlorocyclohexane spectrum are similar, except that signal I is not so well resolved into the methylene groups adjacent to —CHX— and the remote methylene peaks.

The lack of two main components for the low-field signal in each case at 28° C shows a rapid equilibrium between the two chair forms. The —CHX— proton is rapidly converted between axial and equatorial position. On the cooling of both solutions, a distinct broadening of the low-field signal is observed at −40° C and single broad signal occurs at −55° C. Below this temperature two signals are resolved which do not change in appearance appreciably between −70° C and the solidification point (∼−95° C for the chlorocyclohexane solution and ∼−110° C for the bromocyclohexane solution). Marked changes in the general shape of the high-field signal can be detected as high as −25° C. The bromocyclohexane spectrum at high field shows the appearance of an additional low field peak, but the lack of resolvable structure in chlorocyclohexane makes it difficult to classify the changes which occur here.

The chemical shifts of all signals in these spectra are shown in Table I. The highest field component of signal II in both compounds does not show large variation with temperature. The assignment of spectra at low temperature is based on the previous experience with locked conformations where it is found that the low field signal is due to equatorial protons in the cyclohexane ring (3, p. 390 ff.). In the hexachlorocyclohexanes the downfield shift is approximately 0.2–0.6 p.p.m. in agreement with the measured value for chlorocyclohexane at −70° C and −90° C of 0.65 and 0.58 p.p.m.* The corresponding chemical shift separation for the axial and equatorial —CHBr— polar is 0.68 p.p.m. The peak for the equatorial position is narrower than that for the axial position. This finds a ready explanation in terms of the coupling constants involved. The two environments of the —CHX— proton are set out below.

**The center of the roughly gaussian line shape is used for the chemical shift measurement. In view of the overlapping splitting due to J coupling which is symmetrical about the proton concerned, this is a reasonable assumption.*

TABLE I

Chemical shifts of signals in the spectra of chloro- and bromo-cyclohexane at various temperatures with respect to tetramethylsilane

	Chemical shifts (p.p.m.) from $Si(Me)_4$ signal			
	Chlorocyclohexane			
	I		II	
Temp., °C	Weighted average for equatorial and axial		(i)	(ii)
28	−3.935		−2.31 −2.11	−1.54
−26	−3.890		−1.79	−1.55
	(i)	(ii)		
−70	−4.34	−3.76		−1.58
−91	−4.44	−3.79	−2.09	−1.58
	Equatorial proton	Axial proton	(i) Shoulders appearing due to $-CH_2-$ protons adjacent to —CHX— group	(ii) Remote $-(CH_2)_3-$ group
	Bromocyclohexane			
	(i) and (ii)		(i)	(ii)
28	−4.16		−1.95	−1.55
−29	−4.16		−1.95	−1.55
−75	−4.64,	−3.97	−2.12, −1.84	−1.55
−95	−4.65,	−3.97	−2.16, −1.84	−1.55
−104.5	−4.65,	−3.98	−2.17, −1.84	−1.59
	Equatorial proton	Axial proton	$-(CH_2)-$ protons adjacent to —CHX—	Remote $-(CH_2)_3-$ group

If the —CHX— proton is equatorial then coupling constants to the adjacent $-CH_2-$ groups involve ee and ea protons. If the —CHX— proton is axial, aa and ae coupling constants are involved. A comparison of typical coupling constants in some acetylated sugars (3, p. 397) shows that e–e coupling is generally less than 3 c.p.s. and e–a coupling is about 2–5 c.p.s. while a–a coupling is 6–12 c.p.s. The much broader signal for the axial proton is consistent with a larger a–a coupling constant and resolution of some structure is possible within the broad peak (Figs. 2*a* and 2*b*). The spectra at low field

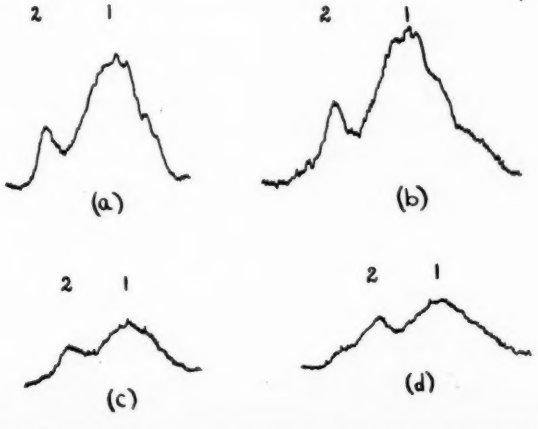


FIG. 2. Spectra of low-field —CHX— proton resonance signal at two temperatures. (*a*) Bromocyclohexane −88° C. (*b*) Bromocyclohexane −97° C. (*c*) Chlorocyclohexane −80° C. (*d*) Chlorocyclohexane −92.5° C.

consist of the 'X' part of an A_2B_2X spectrum in the terminology of Bernstein, Schneider, and Pople (14).

Populations of the equatorial and axial —CHX— proton in equilibrium are given in Table II at various temperatures. The populations were obtained by weighing the peaks after they were traced on to thick graph paper. At each temperature four independent spectra were selected and each traced five times. The total weight of five tracings of each spectrum was then obtained. The smaller peak was then cut off and the remaining axial peak weighed.

TABLE II

Measured populations of equatorial halogens in two cyclohexane derivatives at various temperatures between −80° C and −104° C

Monobromocyclohexane, 55.4 mole% in CS_2		Monochlorocyclohexane, 54.5 mole% in CS_2	
Temp., °C	Equatorial Br, mole%	Temp., °C	Equatorial Cl, mole%
−86	78.2	−80	77.0
−87	79.0	−82	79.9
−89	85.4	−91	71.3
−89	73.2	−92	79.5
−89	81.6	−92	80.9
−89	75.6	−93	72.0
−97	84.8	−93	79.8
−97	81.0		
−97	82.7		
−98	83.4		
−98	83.7		
−98	86.6		
−99	82.9		
−103	80.4		
−103	79.0		
−103	76.4		
−103	79.6		
−103	81.4		
−104	79.3		
−104	83.2		
−104	84.1		

NOTE: Mean values of populations: equatorial bromine = 82.0%; equatorial chlorine = 77.2%.

DISCUSSION

The results in Table II show no significant changes in population of the equatorial halogen form over an ~25° temperature range near −100° C. The energy content of the two equilibrium forms is almost the same. We may calculate a 'ΔF' free energy change from the equation $\Delta F = -RT \log_e K$. The values of the ΔF function are

[1] $$\Delta F_{ea} = -RT \log_e \frac{[\text{equatorial } 'X']}{[\text{axial } 'X']},$$

$$\Delta F_{ea} = -511 \text{ cal/mole for } C_6H_{11}Br,$$

$$\Delta F_{ea} = -406 \text{ cal/mole for } C_6H_{11}Cl.$$

We may assume $\Delta S_{ea} = 0$. Thus the Gibbs free energy function becomes equal to the enthalpy change ΔH_{ea}, which in turn is almost identical with the energy content function at constant volume ΔE_{ea}. Thus the free energy obtained from experiments can

be equated to changes in the energy in going from one conformation to the other. It is to be noted that such small changes, in energy content observed, give negligible changes in population over a 30° C range so that experimental measurement does not detect them.

The populations of the axial chloro and bromo conformers can be obtained at room temperature by using the weighted average of the chemical shifts for axial and equatorial —CHX— form

$$\delta_M = x_a\delta_a + x_e\delta_e. \quad [2]$$

x_a and x_e are mole fractions of chlorocyclohexane in axial and equatorial form. δ_M is the measured chemical shift. δ_a and δ_e are the chemical shifts of axial and equatorial —CHX— proton available from low-temperature measurements.

The mole% equatorial halogen is computed to be 73.2% in bromocyclohexane and 77.4% in chlorocyclohexane. Since the chemical shift measurements of broad peaks are much less accurate than the intensity measurement at low temperature these estimates are subject to large errors.

The measurement of rates of interconversion between conformers using the standard techniques of n.m.r. (8, 9, 10) is not possible in this case because of complex spectra; the line width parameter T_2 cannot be measured. At the coalescence temperature (10) for axial and equatorial —CHX— protons the fine structure is averaged and it is legitimate to calculate an average lifetime τ_c. We have the equation

$$\tau_c \text{ (sec)} = \frac{1}{\sqrt{2}.\ \pi.\ \Delta\nu} \quad [3]$$

where $\Delta\nu$, frequency separation of resolved signals at low temperature (c.p.s.), is equal to 8.9×10^{-3} sec for bromocyclohexane. The coalescence temperature observed was ~218° K in both cases. If we assume a normal frequency factor of $kT/h \simeq \nu_0$ for the conversion process we may calculate an activation energy from

$$\frac{1}{2\tau_c} = 4.36\times10^{12} \exp\left(\frac{\Delta E}{436}\right). \quad [4]$$

The activation energy for the conversion is found to be 10,850 calories, a value subject to large error. Studies are at present being made on substituted cyclohexane having protons which are not coupled to the methylene protons in the ring and more reliable energy barriers will be available.

Although no accurate calculations are available, the small population of the axial halogen form is considered to be due mainly to van der Waals repulsive forces with the 3.5 axial hydrogen atoms (4, 15). Such a calculation is made here in an attempt to estimate the potential energy difference between the axial and equatorial form based on the Lennard-Jones equation (16).

The values ϵ/k = 277° K and 295° K (17) for the bromine–hydrogen and chlorine–hydrogen interactions are used; k is the Boltzmann constant.

Let $$\sigma = \frac{r \text{ van der Waals}}{2^{1/6}} = 2.55 \text{ Å and } 2.42 \text{ Å}$$

for the bromo and chloro compounds respectively. The calculation considers two repulsion terms of the halogen with a hydrogen atom and the van der Waals interaction of this

halogen in the equatorial position with atoms within the van der Waals interaction sphere.

Interatomic distances in the two conformations are obtained from the work of Hassel and co-workers (18). The calculation neglects distortion of halogen or hydrogens from true axial direction in the axial form. Values of 338.5 calories and 50.5 calories are obtained for bromocyclohexane and chlorocyclohexane respectively. The former represents fair agreement for a crude calculation but the latter, although in the correct order of the measured ΔE_{ea}'s, is only of the correct order of magnitude.

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NUCLEAR SHIELDING PARAMETERS FOR PROTONS IN HYDROGEN BONDS

II. CORRELATION OF CHEMICAL SHIFTS IN INTRAMOLECULAR HYDROGEN BONDS WITH INFRARED STRETCHING FREQUENCIES[1]

L. W. Reeves, E. A. Allan, and K. O. Strømme

ABSTRACT

Nuclear shielding parameters have been obtained for 24 intramolecularly hydrogen-bonded phenols and naphthols. The shielding parameters are corrected for large diamagnetic anisotropies and a value $\Delta\sigma_{OH}$ obtained which represents the change in shielding parameter in parts per million with reference to the infinite dilution chemical shift of phenol, α-naphthol, or β-naphthol. These values of $\Delta\sigma_{OH}$ are approximately proportional to the change $\Delta\nu_{OH}$ in the OH stretching frequency on formation of the hydrogen bond.

The present study follows from a recent publication on nuclear shielding parameters for the strongly hydrogen-bonded proton (1) in orthohydroxy azo compounds. It was shown that the corrected chemical shift correlated well with the known acidity of phenol and β-naphthol and hence with the hydrogen-bonding strength, in a series of compounds of quite similar structure. A recognition of the fact that strong intramolecular hydrogen bonds are very little dissociated in dilute solution enables chemical shifts for a completely formed hydrogen bond to be quoted (2, 3). The study of strong dimer bonds (carboxylic acids) and intramolecular hydrogen bonds, both strong and weak, is extended in this present work in an attempt to make a broad and cross correlation with the shift of the infrared —OH stretching frequency. Many of the compounds studied might be expected to deviate from such a correlation (4) and deviations will be examined in detail in future work. Correlations between the frequency shift of the A—H stretch and heat of formation of the hydrogen bond, internuclear distances X—H- -Y across the bond in series of compounds where X and Y are the same atoms, half width of the A—H stretch in the hydrogen bond, and the intensity change ΔB on formation of a hydrogen bond are now well established (4). It is important therefore to obtain a link between infrared and n.m.r. methods so that hydrogen bond strengths might be indicated by either technique. Such a link has been suggested by earlier work (5, 6) where, however, the consideration of a completely formed hydrogen bond is in doubt since intermolecular hydrogen bonds were studied. There has also been a limited amount of study on chelated enolic hydrogen bonds by Försen and Nilsson (7).

EXPERIMENTAL

Some 26 commercially available substances known to contain intramolecular hydrogen bonds of the type —O—H- -Y were purified by fractional distillation or recrystallization from solvents such as cyclohexane, chloroform, ether, alcohol, or acetone depending on which solvent was the most suitable for the particular compound.

Chemical shifts for the intramolecularly hydrogen-bonded proton were measured. The solvent system used was redistilled carbon tetrachloride stored over P_2O_5 to which was added ½ mole% cyclohexane. The cyclohexane served as an internal reference signal and its concentration was accounted for in the calculation of the mole fraction. When the substances under study were liquid, the concentration was made accurately to

[1]*Manuscript received April 11, 1960.*
Contribution from the Chemistry Department, University of British Columbia, Vancouver 8, British Columbia.

2.0±0.2 mole%. Many of the solid substances were not sufficiently soluble to achieve this concentration and were therefore studied at a concentration of 1±0.2 mole%. A smaller number of substances were not even soluble to 1 mole%; in some of these it was possible to heat the 1% mixture in the sealed evacuated 5-mm sample tubes in hot water and achieve a supersaturated solution of 1 mole% on cooling. In a few compounds the —OH proton resonance was detected in solutions in contact with solid material with a concentration approaching the limit for detection (approximately 0.3 mole%).

A Varian V4300 n.m.r. spectrometer detecting proton resonance at a frequency of 40 Mc/sec was used for this work. The chemical shifts were measured with respect to the signal of cyclohexane using the side band technique. Each chemical shift measurement was counted with a Hewlitt Packard electronic counter model 524B. Chemical shifts of the order of 200–500 c.p.s. were reproducible to ±1.5 c.p.s. using independently prepared solutions. When less than 200 c.p.s. was measured, the chemical shifts had a maximum deviation ±0.75 c.p.s. and did not vary over the concentration range 1 to 2 mole% for a given compound.

Some compounds were examined on a Perkin–Elmer model 21 infrared spectrometer to determine the frequency shift from free to hydrogen-bonded —O—H '$\Delta\nu_{OH}$'. The wave number calibration of the infrared spectrometer was assured by measurements of atmospheric water vapor in the region under study, with single-beam operation. In many cases literature values were available for $\Delta\nu_{OH}$ and agreement with the present measurements was satisfactory. The most precise measurements come from the work of Baker and co-workers (8, 9). These authors used a differential method for detecting frequency shifts in some compounds which gives values of $\Delta\nu_{OH}$ approaching an estimated error of ±1 cm^{-1}. The method of direct reading of the calibrated chart paper can only be accurate to ±3 cm^{-1}, which was the limit of error in the present study. Measurements taken from the literature are acknowledged and compared where they overlap with our results in Table I.

RESULTS AND DISCUSSION

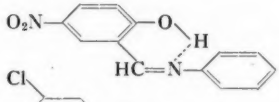
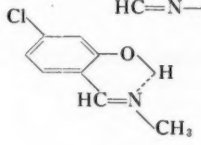
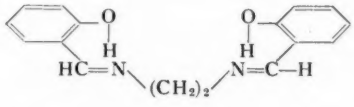
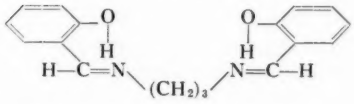
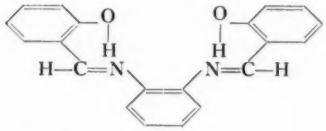
Correlations with Infrared Measurements

The compounds studied contain at least one benzene ring and are of the phenolic —O—H type. The measured chemical shifts have been corrected, in order to reduce their shielding parameters to a common basis. These corrections are for the low-field shift due to the diamagnetic anisotropy of the π electron system (10). The bond distances and angles, where available, were taken from the literature (11), and in cases where the actual distances in the compound under study were not available, a reasonably precise molecular structure could be constructed by considering known structures of related molecules. The correction for the large anisotropy is not precise enough to require the position of the proton in the molecule more accurately than can be estimated in this way. The corrections for the present series of compounds are almost the same in all cases.

The position of the phenol —O—H peak is −2.80 p.p.m. (−112 c.p.s.) with respect to cyclohexane at infinite dilution, a figure in agreement with the previously reported value (12). Corresponding figures obtained in this study for α- and β-naphthols are −3.78 p.p.m. (−151 c.p.s.) and −3.49 p.p.m. (−139.5 c.p.s.).

On the basis of differences in magnetic anisotropy the difference in chemical shift

TABLE I

	Compound	$(\sigma_{cC_6H_{12}} - \sigma_{\text{-O-H}})$, p.p.m.	$(\sigma_{cC_6H_{12}} - \sigma_{\text{-OH}})$ corrected, p.p.m.‡	$\Delta\sigma_{\text{-OH}}$ corrected, p.p.m.‡	$\Delta\nu_{OH}$, cm^{-1}
1.	Salicylaldehyde	−9.51	−8.83	6.71	471*
2.	5-Nitrosalicylaldehyde	−10.05	−9.37	8.70	500*
3.	5-Bromosalicylaldehyde	−9.38	−8.70	6.58	454*
4.	*o*-Nitrophenol	−9.14	−8.46	6.34	346§, 364*
5.	2,4-Dinitrophenol	−9.64	−8.96	6.84	388§
6.	Methylsalicylate	−9.12	−8.44	6.32	395*
7.	*o*-Bromobenzoic acid	−11.25	−10.70	8.58	570*, 96†
8.	*o*-Chlorophenol	−3.97	−3.27	1.17	61§
9.	2,4-Dichlorophenol	−3.99	−3.31	1.19	63§
10.	2,4,6-Trichlorophenol	−4.20	−3.52	1.40	75§
11.	*o*-Bromophenol	−3.95	−3.27	1.15	92*
12.	2,4-Dibromophenol	−3.98	−3.30	1.18	83§
13.	*o*-Iodophenol	−4.13	−3.45	1.33	105§
14.	*o*-Methoxyphenol	−3.98	−3.30	1.18	60§
15.	2,6-Dimethoxyphenol	−3.82	−3.14	1.02	56§
16.	*o*-Allylphenol	−3.87	−3.19	1.07	63§
17.	*o*-Cresol	−3.12	−2.44	0.32	−8*
18.		−12.49	−11.54	9.42	Overlaps —C—H stretch <3000 cm^{-1}
19.		−11.27	−10.59	8.47	
20.		−11.13	−10.45	8.33	544*
21.		−11.36	−10.68	8.56	532*
22.		−11.03	−10.08	7.96	531*
23.	1-Nitro-2-naphthol	−10.53	−9.70	7.04	538*
24.	2,4-Dinitro-1-naphthol	−11.20	−10.28	7.42	543*

NOTE: Reference for benzene derivatives is OH proton signal in phenol, while for naphthalene derivatives it is OH in 1- and 2-naphthol respectively.
*Measured in this work. †Infrared frequency of small monomer peak. ‡Corrected for ring currents. §cf. reference 8.

between the isolated phenol and β-naphthol molecules should be 0.2 p.p.m. (10). Table II shows the measured chemical shifts of isolated phenol, α-naphthol, and β-naphthol

TABLE II

Infrared stretching frequencies of monomers and isolation chemical shifts in phenol and in 1- and 2-naphthol

Compound (referred to infinite dilution in CCl_4)	$(\sigma_{c\text{-}C_6H_{12}} - \sigma_{OH\text{-}dil})$, p.p.m.	$(\sigma_{c\text{-}C_6H_{12}} - \sigma_{OH\text{-}dil})$ corrected, p.p.m.	ν_{OH}, cm^{-1}
Phenol	2.80	3.12	3604†
2-Naphthol	3.49	2.66	3601*
1-Naphthol	3.78	2.86	3594*

*Measured in this work. †cf. reference 8.

in column 2. Column 3 shows the corrected isolation chemical shifts for these molecules assuming the —O—H atoms lie in the plane of the molecule in one of two positions (4). It is assumed in the case of β-naphthol that these positions are equally populated. In α-naphthol the cis position to the second aromatic ring is seriously hindered by the α-hydrogen in the second ring so 100% population of the trans position is assumed in the correction.

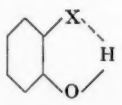
It is interesting to note that the difference in isolation chemical shift of these three molecules after correction is in the same order as the —O—H stretching frequencies. Differences in the corrected chemical shifts are due to changes in the polarity of the —O—H bond and thus reflect the charge density at the proton. It is probable that the acidity increases in the order phenol, β-naphthol, α-naphthol. The wave number measurements of the free —O—H stretch frequency were not accurate enough to indicate the precise correlation between isolation chemical shift and frequency shift.

Figure 1 shows the differences in shielding parameter between the free —O—H signal and the hydrogen-bonded proton plotted against the infrared frequency shift of the —OH stretch on formation of the intramolecular hydrogen bond. The plot shows an approximate linear relationship from the weakest hydrogen bond 2,6-dimethoxyphenol to the strong hydrogen bond in some orthohydroxy Schiff's bases.

In intermolecular hydrogen bonds the heat of formation correlates well with the frequency shift $\Delta\nu_{OH}$ (4). The intensification of the —OH stretch vibrational transition can also be used to indicate the heat of formation and the intensification is more sensitive detection of hydrogen bond formation. In the formation of intramolecular hydrogen bonds, however, only the frequency shift $\Delta\nu_{OH}$ shows normal behavior in correlating with the heat of formation of the bond (4).

Unpublished work of Pimentel and Brown (4) shows that the original linear correlation of Badger and Bauer (14) does not hold but that low-frequency shifts $\Delta\nu_{OH}$ correspond to disproportionately high heats of formation compared with the large values of $\Delta\nu_{OH}$. The systems studied were phenol in a series of bases and therefore relevant to the present study of intramolecularly hydrogen-bonded phenols.

Weak Intramolecular Hydrogen Bonds



In compounds of the type shown above where a weak hydrogen bond is formed as in

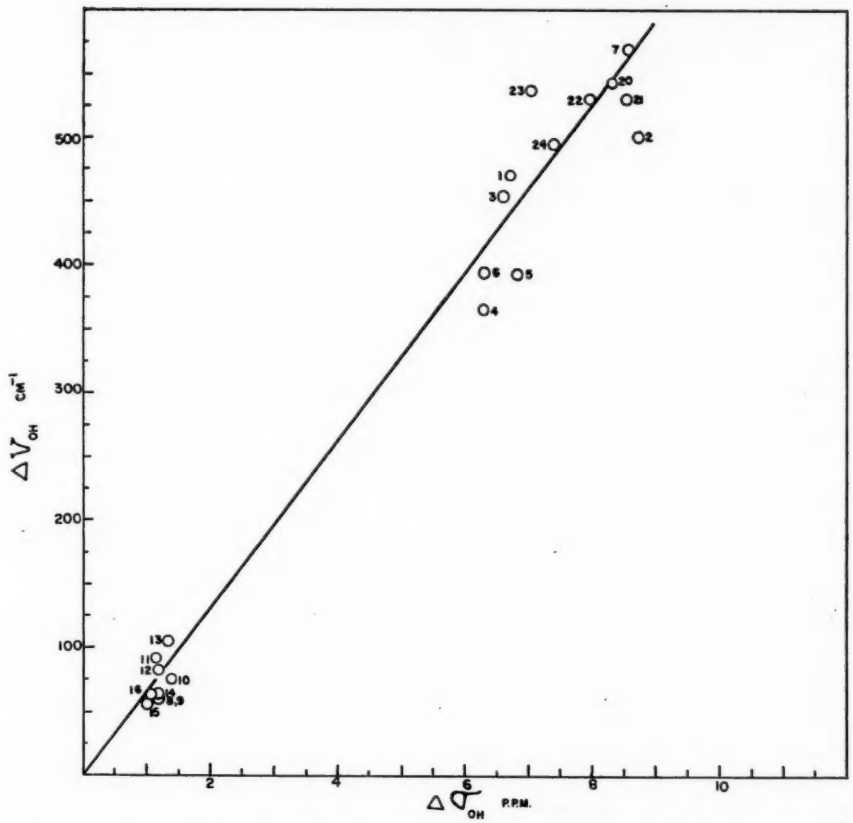


FIG. 1. Plot of $\Delta\nu_{OH}$ versus $\Delta\sigma_{OH}$ for intramolecular hydrogen bonds in phenols and naphthols. Numbered points refer to compounds in Table I.

the case of the halogens or orthomethoxy phenols, the population of the trans position must be taken into account, in order to justify a measurement of the change in shielding parameter on hydrogen bond formation. The careful infrared studies of Baker and co-workers (8, 9, 13) show that the free phenolic —OH stretch is observed as well as the band due to a hydrogen-bonded proton (15). After careful purification of the *o*-halophenols Baker concluded that the cis/trans ratios in *o*-chloro-, *o*-bromo-, and *o*-iodo-phenol were 1/56, 1/38, and 1/13.5 respectively. If we assume the chemical shifts in these compounds to be characteristic of the average environment then there is less than a 10% error in considering the most unfavorable case *o*-iodophenol as a completely cis configuration.

The chemical shift changes on formation of the hydrogen bond are roughly proportional to $\Delta\nu_{OH}$. Baker has shown, however, that hydrogen bond strengths are in the order Cl > Br > F > I (9). In these molecules the O—H- -O angle is distorted at least 85° from a straight line and interorbital repulsion plays a part in the anomalous order of $\Delta\nu_{OH}$ (9).

There are reasons why anomalies should occur in the chemical shift measurement. In

addition to steric factors which favor the order suggested by Baker there is the increasing diamagnetic anisotropy of the C—X bond as X changes from fluorine to iodine. The ratio $\Delta\nu_{OH}/\Delta\sigma_{OH}$ shows that in going from chlorine to iodine we have the values 50.2, 67.8, and 78.8 cm^{-1}/p.p.m., indicating a term which lowers $\Delta\sigma_{OH}$ as we go from chlorine to iodine.

In the weak intramolecular hydrogen bonds there are several exceptions to the correlation of $\Delta\sigma_{OH}$ and $\Delta\nu_{OH}$ with ΔH_f the heat of formation of the bond (9). It must be remembered, however, that this region of the correlation represents a small portion of the total range ($\sim$2.5–3.5 kcal = ΔH_f) of hydrogen bond energies studied (0–8 kcal/mole = ΔH_f).

Further detailed studies of this restricted region are in progress to extract the artifacts peculiar to magnetic properties of these molecules.

ACKNOWLEDGMENT

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OROTIC ACID AND ITS ANALOGUES

PART II. ON THE ALKALINE REARRANGEMENT OF 5-CARBOXYMETHYLIDENEHYDANTOIN[1, 2]

R. DEGHENGHI AND G. DANEAULT

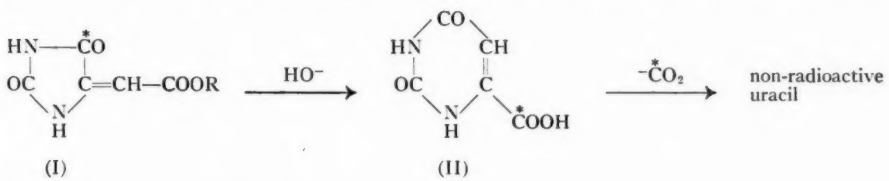
ABSTRACT

The in vitro conversion of 5-carboxymethylidenehydantoin to orotic acid is not likely to occur in vivo since the two substances have antagonistic effects (2). A sulphur analogue of the above hydantoin has been prepared and shown not to undergo alkaline rearrangement to the corresponding metathiazine structure. The mechanism of orotic acid formation is discussed.

Since its discovery in 1905 by Biscaro and Belloni (3), orotic acid (II) has aroused an enormous interest among biochemists (see in particular (4) and (5)) for its importance as a pyrimidine precursor of nucleic acids. A large amount of synthetic work has been done on orotic acid itself (6) and on its analogues (7), but comparatively little attention has been given to the mechanisms of its formation, from a purely chemical point of view.

At first orotic acid was believed to be a seven-membered ring compound (3), but it was soon shown to be uracil-6-carboxylic acid (8, 9). Its synthesis from ethyl oxaloacetate and urea was erroneously believed to give directly ethyl orotate (10); later the ester was proved to be in fact 5-carbethoxymethylidenehydantoin (11). A similar error was made by Biltz and Kramer (12), who studied the alkaline decomposition of a diazoacetate derivative of alloxan to give again the hydantoin and not the six-membered ring ester.

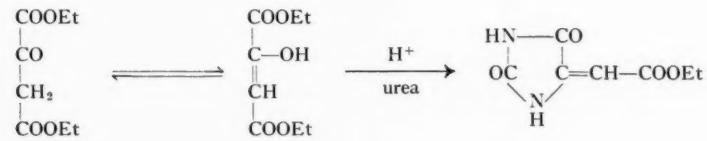
All these facts prompted us to reinvestigate the mechanism of the formation of orotic acid and we chose at first the reaction of ethyl oxaloacetate and urea (III → I), as described by Müller (13) and discussed by Nyc and Mitchell (14). The latter authors postulated a mechanism which involves opening of the hydantoin ring between positions three and four to give the hydantoic acid which closes again to give the pyrimidine. The reaction was followed spectrophotometrically in alkaline solution and it was found that orotic acid is formed previous to acidification. The hydantoic acid was not isolated, though in other unsaturated hydantoins evidence of the open-chain compound has been obtained (14). That an open-chain-ring tautomerism exists in the hydantoin conversion to orotic acid was clearly demonstrated by Langley (15), who was able to show that a C^{14} carbon atom in position four in the hydantoin ring is present as a carboxyl in the orotic acid; this was confirmed by the quantitative recovery of the radioactivity in the CO_2 following decarboxylation (the remaining uracil being inactive). We were interested



(I) (II)

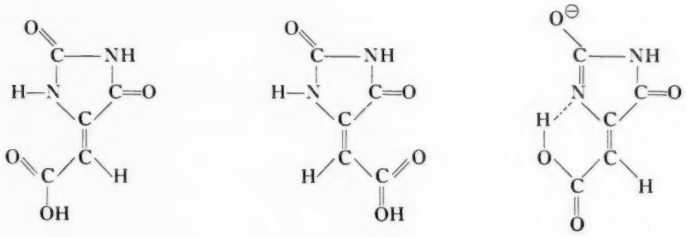
[1]*Manuscript received April 13, 1960.*
Contribution from the Department of Chemistry, Laval University, Quebec, Que. Taken from the thesis submitted by G. Daneault for the degree of Bachelor of Science.
[2]*Part I: See reference 1.*

to investigate (*a*) why the hydantoin is the first intermediate when urea condenses with oxaloacetic ester, (*b*) whether the same kind of ring enlargements were possible with other five-membered ring substances, and (*c*) the isolation of possible intermediates of the above rearrangement (I → II).



(III) (IV) (I)

Oxaloacetic acid and its esters (III) are present in the pure state or in non-polar solvents (16, 17) as the enol form (IV) which may exist in the *trans* (hydroxy-fumaric) or *cis* (hydroxy-maleic) form. It is formally the latter which, after condensation with urea, gives the five-membered ring hydantoin. In fact, one could reasonably expect ethyl orotate only from the *trans* (fumaric) form of the intermediate hydantoic ester (ureidofumaric acid). For the same reason one could expect a cis–trans isomerization to occur previous to or during the alkaline conversion of the hydantoin to orotic acid. We were able to isolate two isomeric forms of the hydantoin (V, VI), one of which has not been described previously, by eluting the potassium salt of I through a column of cation-exchange resin (cf. Experimental). Proof of the structure was obtained by ultraviolet spectra, which showed two peaks at 235 mμ and 294 mμ characteristic of hydantoin structure (cf. 14), by infrared spectra, by the facile conversion of one isomer to the other, and by the fact that both substances gave, upon reduction with sodium amalgam, the same hydantoin-5-acetic acid. Furthermore, we were led to attribute tentatively structure V to the lower melting isomer, on the basis that the increase of the infrared frequency of the carbonyl in position four and the decrease of the C=O frequency in position two, together with the shift of the OH frequency due to hydrogen bonding, are better explained by structures of the type VII. Therefore the complete mechanism

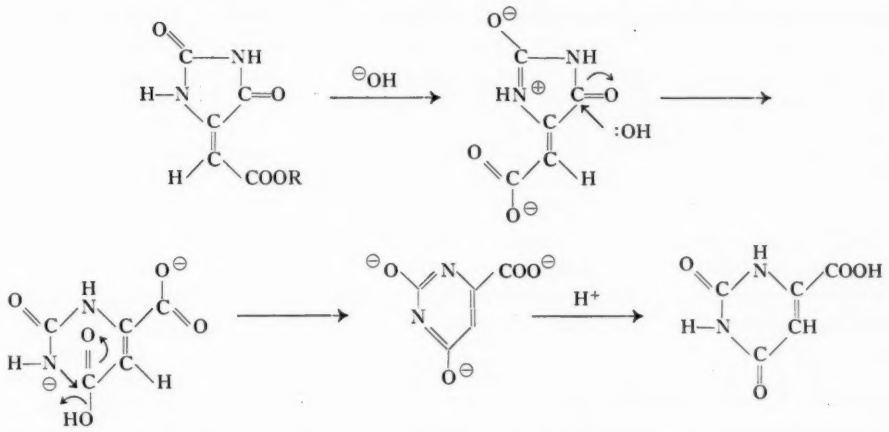


(V) (VI) (VII)

for the alkaline conversion of I to orotic acid may reasonably assume the following form going through the steps: (*a*) cis–trans isomerization,[3] (*b*) alkaline cleavage, and (*c*) aromatization to the more stable pyrimidine structure.[4]

[3]*The reversible addition of the elements of water to the double bond (cf. 19) is a possible mechanism for this isomerization.*

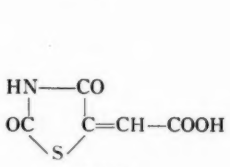
[4]*Orotic acid is essentially a hydroxy pyrimidine in alkaline medium (cf. 18).*



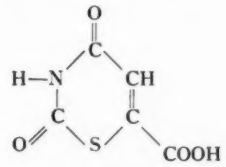
For a parallel experiment with a parent compound, we chose the corresponding pseudo-thio-hydantoin (VIII) which we synthesized in good yields from thiourea and maleic anhydride through the saturated thiazolidinedione-5-acetic acid. The ultraviolet spectrum of the new substance VIII in ethanol resembles closely that of 5-carboxymethylidene-hydantoin (cyclic, five-membered structure) and the infrared spectrum of the corresponding methyl or ethyl esters shows the four expected bands at 1772 cm^{-1} (4-CO), 1735 cm^{-1} (2-CO), 1690 cm^{-1}, and 1620 cm^{-1} (α,β-unsaturated ester). The free acid had a similar spectrum but with two more bands (KBr) at 1785 and 1700 cm^{-1} which could possibly arise from dimerization[5] (20). The free acid has a remarkable stability; it was recovered unchanged after being refluxed with an aqueous solution of $HClO_4$ and, when subjected to the same conditions for the orotic acid alkaline rearrangement, the starting material or its monopotassium salt were recovered again in high yield (cf. Experimental).

The potassium salt[6] by elution through a cationic exchange resin gave, like the above hydantoin, an isomer whose cyclic structure was evidenced by ultraviolet absorption, and which could be converted to the other form by crystallization from boiling water. Both isomers were reduced to the same saturated acid by the action of sodium amalgam.

The fact that in the case of the five-membered ring sulphur compound a ring enlargement to the corresponding metathiazine structure (IX) was not achieved, in spite of



(VIII)



(IX)

(X)

(XI)

[5] *For a discussion about carbonyl band splitting, cf. (26).*
[6] *Cf. (9) for the remarkable stability of potassium orotate.*

the isomerization, may be explained by considering the stability of structures of the type X which could arise from the enolic form of the hydantoin (21). The polarity of the exocyclic double bond (conjugated with the acid) would be strongly opposed in the structure X as it is in the case of 1,2-cyclopentenediones (enol form XI) (22) which, accordingly, exist only in the keto form. With the sulphur compound evidently no enol structure of the type XI (in alkaline solution) would be possible, thus explaining the observed stability of the five-membered thiazolidine ring.

It may be mentioned here that the in vitro conversion of 5-carboxymethylidene-hydantoin to orotic acid is not likely to occur in vivo, since the two substances show antagonistic effects following subcutaneous administration in the rat (2). An account of the biological activity of the sulphur analogue of the above hydantoin will be reported elsewhere.

EXPERIMENTAL[7]

2-Imino-4-oxo-5-thiazolidineacetic Acid

This compound was prepared in 72% yield from thiourea, according to the method of Andreasch (23) reported by Flett and Gardner (24), m.p. 252–253° dec. (reported to decompose above 220°). ν_{max}^{KBr} 3300, 3150 cm^{-1} (OH, NH); 2400 cm^{-1} (broad); 1690, 1645 cm^{-1} (C=N and C=O). Calc. for $C_5H_6N_2O_3S$: C, 34.48%; H, 3.47%; N, 16.09%; S, 18.40%. Found: C, 34.62%; H, 3.71%; N, 16.34%; S, 18.28%.

2,4-Dioxothiazolidine-5-acetic Acid

A solution of 63.237 g of 2-imino-4-oxo-5-thiazolidineacetic acid was refluxed for 2 hours with 630 ml of 20% sulphuric acid. Upon cooling, 42.6 g of 2,4-dioxothiazolidine-5-acetic acid crystallized out, m.p. 168° (lit. 168–169° (25)).

Extraction of the mother liquors with ethyl acetate and evaporation of the solvent *in vacuo* gave another 19.9 g of less pure product, m.p. 163–168°. Total yield 98%. ν_{max}^{KBr} 3200, 3080 cm^{-1} (OH, NH); 1760, 1740 cm^{-1} (4 and 2 C=O); 1700 cm^{-1} (acid C=O).

5-Carboxymethylidene-2,4-dioxothiazolidine (VIII)

A mixture of 26.25 g of 2,4-dioxothiazolidine-5-acetic acid in 90 ml of acetic acid containing 8.2 ml of bromine was refluxed for 2 hours. Hydrogen bromide was evolved and from the clear solution, upon cooling, the product crystallized out in 60% yield. The crystals were collected on a filter, washed with a little acetic acid and water, m.p. 248° dec. A sample was recrystallized from water three times for analysis; needles, m.p. 252–253° dec. Calc. for $C_5H_3NO_4S$: C, 34.68%; H, 1.75%; N, 8.09%; S, 18.51%. Found: C, 34.80%; H, 1.92%; N, 7.98%; S, 18.30%. λ_{max}^{EtOH} 236 mμ (ε 4360), 302 mμ (ε 6320). ν_{max}^{KBr} 3240, 3080, 2950 (OH, NH); 1775, 1758 cm^{-1} and 1725, 1700 cm^{-1} (split carbonyls); 1670 and 1622 cm^{-1} (α,β-unsaturated acid).

The *methyl ester* was prepared by refluxing the acid (0.5 g) in 10 ml of methanol and 0.5 ml of sulphuric acid for 1 hour. Dilution with water gave 5-carbomethoxymethylidene-2,4-dioxothiazolidine, m.p. 160–162°, in 80% yield. The product was recrystallized from water for analysis, m.p. 163–164°. Calc. for $C_6H_5O_4NS$: C, 38.49%; H, 2.69%; N, 7.48%; S, 17.13%. Found: C, 38.24%; H, 2.89%; N, 7.66%; S, 17.21%. ν_{max}^{KBr} 3250 cm^{-1} (NH); 1770 and 1732 cm^{-1} (4 and 2 C=O); 1690 and 1618 cm^{-1} (α,β-unsaturated CO).

The *ethyl ester* was prepared in an analogous manner, m.p. 128–130°, λ_{max}^{EtOH} 240 mμ

[7]*Melting points, taken in evacuated capillaries, are corrected. Analyses by Schwarzkopf Microanalytical Laboratories, Woodside, N.Y. The infrared spectra were taken with a Beckman IR-4 spectrophotometer* (NaCl *optics).*

(ϵ 3940), 305 mμ (ϵ 6000). ν_{max}^{KBr} 3300 (NH); 1770 and 1720 cm^{-1} (4- and 2- CO); 1690 and 1630 cm^{-1} (α,β-unsaturated CO).

Stability of 5-Carboxymethylidene-2,4-dioxothiazolidine in Acidic Conditions

To a solution of VIII, 0.449 g in 20 ml of water, was added 2 ml of 72% perchloric acid and the whole was refluxed for 30 minutes. Upon cooling, 0.3 g of the starting material (VIII), m.p. 248° dec., crystallized out. The product was identified by infrared spectra and mixed melting point with authentic material.

Stability of 5-Carboxymethylidene-2,4-dioxothiazolidine in Alkaline Conditions

A solution of VIII, 0.693 g in 15 ml of 1 *N* KOH, was kept 30 minutes at 100°. After cooling, 8 ml of 1 *N* HCl was added. The monopotassium salt of VIII crystallized out; needles, m.p. 270° dec. (with previous browning). Calc. for $C_5H_2NO_4SK$: C, 28.43%; H, 0.95%; N, 6.63%. Found: C, 28.55%; H, 1.02%; N, 6.69%. λ_{max}^{EtOH} 235 mμ (ϵ 1330), 298 mμ (ϵ 2600). ν_{max}^{KBr} 2550 (broad); 1730, 1690, 1630, 1580 cm^{-1} (4- and 2-CO, α,β-unsaturated ionized acid). The same salt (infrared and mixed melting point comparison) was obtained when the unsaturated acid (VIII) was neutralized with 1 *N* KOH in the cold, followed by concentration of the solution *in vacuo*.

Lower Melting 5-Carboxymethylidene-2,4-dioxothiazolidine

A solution of 0.782 of the potassium salt of VIII in water was passed through a column of cation exchange resin Amberlite I.R.-120 (H). The aqueous fractions were extracted with ethyl acetate and the organic solvent was evaporated *in vacuo* to give a crystalline residue (0.7 g), m.p. 200–205° dec. ν_{max}^{KBr} 3400, 3220, 3080 (OH, NH); 1768 and 1722 (4 and 2 C=O); 1690 (shoulder); 1670 and 1615 cm^{-1} (α,β-unsaturated CO). λ_{max}^{EtOH} 240 mμ (ϵ 3840), 303 mμ (ϵ 4000). When this substance was crystallized from water, the original higher melting isomer was obtained, and identified by infrared and mixed melting-point determinations.

Reduction of 5-Carboxymethylidene-2,4-dioxothiazolidine

To a suspension of 0.759 g of VIII in 25 ml of water, stirred in an ice-water bath in a CO_2 atmosphere, 25 g of 2% sodium amalgam was added piecemeal.

After 2 hours the clear solution was decanted from the mercury, acidified with 20% sulphuric acid, and extracted with ethyl acetate. Evaporation of the solvent gave 217 mg of 2,4-dioxothiazolidine-5-acetic acid, identified by mixed melting point and infrared determination with authentic material.

When the *lower melting isomer* was reduced in a similar way, the same product, 2,4-dioxothiazolidine-5-acetic acid, was obtained in similar yield.

5-Carbethoxymethylidenehydantoin (I, R = Et)

This substance was prepared according to the method of Müller (13) from ethyl-oxaloacetate and urea. The yields were in the 10–20% range, m.p. 184–186° (reported 188°). λ_{max}^{EtOH} 235 mμ (ϵ 7920), 300 mμ (ϵ 7500), ν_{max}^{KBr} 3250 (NH); 1800, 1760 (4- and 2-CO); 1695 (C=N), 1680, 1640 cm^{-1} (shoulder) (α,β-unsaturated CO).

5-Carboxymethylidenehydantoin (I, R = H)

This acid was prepared by careful hydrolysis of the corresponding ester (I, R = Et), 184 mg, dissolved in 2.3 ml of ethanol, by addition of 0.1 g KOH in 0.4 ml water. The yellow solid formed was dissolved in 5.6 ml of water. Acidification with concentrated HCl precipitated the product, m.p. above 300° with gradual decomposition. λ_{max}^{EtOH} 237 mμ

(ϵ 7800), 300 mμ (13000). ν_{max}^{KBr} 3550, 3300, 3100 cm^{-1} (OH, NH); 1790 and 1745 cm^{-1} (4 and 2 C=O); 1680 (broad); 1638 cm^{-1} (shoulder) (α,β-unsaturated C=O). A lower melting isomer was obtained by dissolving the above acid in 1 *N* KOH and by eluting the slightly alkaline solution through a column of Amberlite I.R.-120 (H) (cation exchange resin).

The aqueous fractions combined were extracted with ethyl acetate, the solvent removed *in vacuo* to give an acid, m.p. 250° dec. ν_{max}^{KBr} 3250 (broad); 2995 (bonded OH, NH); 1800 cm^{-1} (4-CO in β,γ-unsaturated ring), 1725 (2 CO or C=N), 1690 cm^{-1} (acid CO). λ_{max}^{EtOH} 235 mμ (ϵ 5150), 294 mμ (ϵ 8850). When this substance was recrystallized from water, the higher melting isomer was obtained (infrared spectra identical). Both isomers gave, upon reduction with Na/Hg (compare the above described procedure), the same 5-hydantoin acetic acid, m.p. 212° dec. (lit. 214° (14)).

Orotic Acid

This substance was obtained by the alkaline rearrangement of 5-carboxymethylidenehydantoin according to Nyc and Mitchell (14), m.p. 320° dec. ν_{max}^{KBr} 3580, 3150, 3050 cm^{-1} (OH, NH); 1740 (shoulder); 1715 and 1680 cm^{-1} (broad) (C=N, C=O). The infrared spectrum was identical with that of an authentic sample.

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THE THERMAL DECOMPOSITION OF PHENYLACETIC ACID[1]

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ABSTRACT

The thermal decomposition of phenylacetic acid was investigated by the toluene-carrier technique over the temperature range 587 to 722° C. The products of the pyrolysis were carbon dioxide, carbon monoxide, hydrogen, methane, dibenzyl, and phenylketene. From the kinetics of the decomposition it was concluded that the reaction

$$C_6H_5CH_2COOH \rightarrow C_6H_5CH_2 + COOH$$

was a homogeneous, first-order process and that the rate constant of this dissociation step was represented by the expression $k = 8\times10^{12}.e^{-55{,}000/RT}$ sec^{-1}. The activation energy of this reaction may be identified with $D(C_6H_5CH_2—COOH)$. The possible reactions of carboxyl radicals are discussed.

INTRODUCTION

Due to the resonance stabilization of the benzyl radical, the dissociation energies of the bonds $C_6H_5CH_2—R$ are smaller than the corresponding bond dissociation energies in related compounds of the type $CH_3—R$, where R is a radical or atom (1, 2). Therefore, compounds of the type $C_6H_5CH_2—R$ have proved particularly suitable for determination of these dissociation energies and of the heats of formation of radicals R using the toluene-carrier technique (3, 4, 5).

In the present study the thermal decomposition of phenylacetic acid in a stream of toluene was investigated for the purpose of producing carboxyl radicals by the rupture of the $C_6H_5.CH_2—COOH$ bond and for studying their subsequent reactions. Furthermore, it was anticipated that from the kinetics of the pyrolysis one might be able to derive the dissociation energy of this bond.

EXPERIMENTAL

Materials

Phenylacetic acid (m.p. 77° C) was obtained from Eastman Kodak Co. and used without further purification. The toluene was generously supplied by Gulf Oil Corporation, Pittsburgh, Pa. It was stirred for 24 hours with dilute sodium hydroxide, dried, and distilled through an efficient column. The fraction boiling at 110.5° C was used in this study.

Apparatus and Procedure

The pyrolysis was studied in a vacuum flow system similar to those used in previous investigations employing the toluene-carrier technique (6). The reaction vessel was 26 mm in diameter and 20 cm long, with a volume of about 200 cc. It was provided with a central thermocouple well to enable one to measure the temperature along the length of the whole reaction vessel. The gradient along this length was not more than 4 degrees and fell sharply on either side. Adequate preheating was ensured by allowing the gas to pass through a length of tubing intermediate in size between that of the reaction vessel and the inlet lead. The temperature of the furnace surrounding the reaction vessel was regulated within ±2° with a Thermo Electric regulator.

The surface dependence of the reaction was examined by using a similar reaction vessel

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packed with silica wool, supplied by Microchemical Specialties, Berkeley, Calif. The average diameter of the silica fibers was 3×10^{-3} cm and the average length was 15 cm. The surface-to-volume ratio of this vessel was about 27 times that of the unpacked vessel. The decomposition of phenylacetic acid was studied over the temperature range 587 to 772° C.

Known weights of the acids were introduced into a U-tube in the form of pellets. This method proved convenient and accurate and avoided errors due to handling losses of the solid powder. One arm of the U-tube was connected with the toluene reservoir and the other led directly to the furnace. The U-tube was heated by a removable bath containing dibutyl phthalate for the lower temperatures and Fisher bath wax for the higher temperatures. The temperature of the bath was maintained constant with a Fenwall thermoregulator.

The tubes leading to and from the furnace were heated electrically with Nichrome wire to prevent condensation of the acid and products. The pressure of toluene and reactant during an experiment was maintained constant with a manostat and measured with a mercury manometer, read with a cathetometer to an accuracy of ±0.01 mm. Variations in pressure caused the mercury in the sloping arm of the manostat to make and break contact with a platinum wire connected to a suitable relay, which in turn controlled the heater of the water bath surrounding the toluene.

The products were fractionated through a series of three traps maintained at −5° C, −78° C, and −188° C. The non-condensable gases were removed by two diffusion pumps connected in series and collected into calibrated bulbs. The undecomposed acid and dibenzyl were trapped together at −5° C and weighed. The recovered acid was titrated and the yield of dibenzyl was calculated by difference. The dibenzyl collected from a few experiments was purified by recrystallization and mixed with a pure sample; its mixed melting point was 52° C. Carbon dioxide was measured in a gas burette and identified by its vapor pressure – temperature curve. Samples of the non-condensable gases were analyzed by combustion over copper oxide at 300° C.

RESULTS

The main product of the decomposition was carbon dioxide, with much smaller amounts of carbon monoxide. Hydrogen and methane were formed in amounts varying between 20 and 60% of the carbon dioxide, and the ratio of hydrogen to methane varied from 2.0 to 3.5. The yield of dibenzyl approached that of the non-condensable gases. Table I shows the main products and the amounts of acid decomposed under different experimental conditions. The experiments marked with letter P refer to those done in the packed vessel.

The carbon dioxide and carbon monoxide did not account for all the acid decomposed. It was found that the decomposition also yielded phenylketene,* a yellow compound which was trapped with the toluene. The ketene reacted with water to give phenylacetic acid which was titrated with sodium hydroxide. However, the amounts of phenylketene recovered in this way did not account completely for the acid deficit. Phenylketene has been reported to be a very reactive compound which rapidly dimerizes (8) and this might have prevented its quantitative conversion to the acid. The production of phenylketene was considerably increased in the packed reaction vessel, and this mode of decomposition

**Reaction of this compound with aniline gave the anilide of phenylacetic acid, which was identified by a mixed melting point determination with a pure sample prepared according to the method described by Shriner and Fuson (7). The melting point of the mixed sample was 114 to 115° C.*

TABLE I

Products of the decomposition of phenylacetic acid

Expt. No.	T, °A	Total pressure, mm Hg	P_a/P_t*	t, sec	Acid decomp., mmoles	% decomp.	CO_2, mmoles	Non-condensable gases, mmoles	% CH_4	% CO	% H_2	Dibenzyl / H_2+CH_4+CO	H_2 / CH_4	k (sec^{-1})	k_1 (sec^{-1})
123	860	12.45	.038	1.03	.57	19.6	.191	.038	27.4	12.3	60.2	.87	2.2	.21	.07
124	870.5	12.8	.036	.98	.57	21.0	.234	.059	28.9	9.1	62.0	.73	2.1	.24	.10
119	876	10.8	.027	.35	.375	11.4	.1415	.020	27.8	8.8	63.4		2.3	.35	.13
136P	876	11.5	.010	.42	2.26	73.0	.128	.0265	12.6	52.5	34.9		2.8	2.98	.18
118	877	11.3	.030	.41	.36	12.0	.147	.024	28.0	8.9	63.2	.75	2.3	.31	.12
98	879	6.3	.016	.37	.40	19.2	.119	.021	28.8	13.3	57.9	.71	2.0	.58	.17
120	881	12.0	.024	.36	.32	12.8	.131	.021	28.3	7.3	64.4		2.3	.38	.15
96	881	12.1	.017	.49	.43	21.5	.139	.028	30.6	11.2	58.2	.61	1.9	.49	.16
97	881	5.5	.033	.85	.61	30.6	.232	.068	27.2	11.1	61.7	.84	2.3	.43	.17
131P	881.5	12.2	.013	.345	1.38	57.0	.123	.018	15.7	45.0	39.3		2.5	2.45	.22
125	882.5	13.1	.033	.95	.67	24.5	.340	.1075	29.1	7.7	63.2	.81	2.2	.30	.15
115	891.5	4.7	.029	.42	.38	16.9	.1915	.043	27.1	8.3	64.6	.67	2.4	.44	.22
116	892	12.7	.019	.39	.43	17.6	.201	.201	29.5	6.2	64.4	.80	2.2	.50	.23
135P	892	12.2	.010	.39	1.67	77.5	.123	.031	14.2	51.4	34.4		2.4	3.88	.31
126	904	11.6	.013	.975	.87	47.5	.417	.208	28.1	10.2	61.7	.84	2.2	.66	.34
109	904	11.4	.012	.31	.31	18.1	.175	.041	30.9	5.6	63.7	1.00	2.1	.65	.36
111	904	11.25	.013	.29	.32	16.1	.213	.041	31.4	4.6	64.0	.66	2.0	.60	.40
110	904.5	11.4	.016	.41	.34	19.8	.230	.067	30.8	4.3	64.9	.79	2.1	.54	.36
82	914	11.2	.015	.345	.39	25.5	.224	.0765	30.2	9.6	60.2	.82	2.0	.85	.50
134P	920	11.5	.0083	.33	.31	20.6	.139	.061	15.5	50.5	33.9		2.2	6.18	.78
94	926	12.9	.0094	.30	.69	44.1	.253	.090	29.4	9.2	61.4	.76	2.1	1.58	.69
93	926	5.3	.031	.47	.60	47.6	.312	.142	22.4	12.5	65.1	.89	2.9	1.36	.74
127	928	10.2	.022	1.05	1.07	69.8	.569	.4115	24.1	12.9	63.0	.81	2.6	1.13	.65
117	931.5	10.6	.038	.33	.69	32.3	.4605	.193	25.1	5.8	69.1	.88	2.8	1.20	.81
100	938	10.3	.011	.27	.73	43.7	.316	.143	27.8	10.8	61.4	.74	2.2	2.12	.95
101	938.5	10.9	.018	.56	.96	67.0	.477	.307	25.6	13.4	61.0	.79	2.4	1.96	1.05
102	939	11.0	.031	.30	.45	30.4	.388	.159	25.8	2.3	71.9	.54	2.8	1.20	1.03
103	939.5	10.7	.013	.35	.55	34.6	.454	.218	28.6	2.7	68.6	.86	2.4	1.22	1.00
133P	940	11.75	.0081	.30	.48	29.3	.211	.129	16.3	48.4	35.3		2.2	7.83	1.42
83	947	11.3	.011	.28	.72	46.5	.459	.241	26.8	9.1	64.1	.76	2.4	2.24	1.47
84	950	10.0	.038	.34	.93	55.6	.601	.305	20.8	10.8	68.4	.95	3.3	2.44	1.64
112	956	4.7	.023	.40	.87	60.9	.629	.356	20.2	10.1	69.8	.89	3.5	2.37	1.80
113	957	11.1	.020	.35	.73	55.5	.554	.3195	24.0	6.6	69.4	.86	2.9	2.32	1.81
114	957	11.4	.016	.28	.67	50.0	.453	.256	25.4	5.9	68.6	.80	2.7	2.33	1.69
93	960.5	11.1	.015	.31	.80	64.4	.517	.315	23.5	11.9	64.6	.80	2.7	3.36	2.28
106	968	4.9	.031	.39	.61	69.1	.515	.286	18.3	7.4	74.4	.91	4.1	2.94	2.60
108	968	11.0	.010	.28	.54	57.4	.4155	.264	25.4	5.9	68.6	.82	2.7	3.05	2.41
107	968.5	10.9	.012	.35	.55	65.2	.461	.295	24.1	6.1	69.8	.83	2.9	3.00	2.64
105	980	10.7	.011	.27	.53	65.7	.462	.303	24.0	5.2	70.7	.80	2.9	4.00	3.55
104	980.5	11.1	.012	.31	.63	70.6	.610	.3725	23.7	4.3	71.9	.83	3.0	4.07	3.91
99	986	6.8	.022	.35	.74	87.8	.507	.371	16.7	20.4	63.0	.77	3.8	5.90	4.63
92	990	6.6	.028	.35	.74	88.7	.581	.388	15.1	19.1	65.8	.77	4.4	6.32	5.52
91	991	11.0	.015	.29	.685	85.9	.501	.383	20.3	17.3	62.3	.72	3.1	6.88	5.61
121	995.5	11.1	.014	.285	.91	85.2	.709	.542	19.8	13.1	67.1	.79	3.4	6.77	5.74
122	995.5	11.1	.012	.26	.86	85.6	.660	.496	20.7	12.4	67.0		3.2	6.60	5.49

*P_a and P_t refer to partial pressures of phenylacetic acid and toluene, respectively.

may be partly heterogeneous. The percentage carbon monoxide in the non-condensable gases also increased to about 50% although the total production of non-condensable gases was only very slightly increased (Table II).

In one experiment the contents of the dry ice – acetone trap were tested for the presence of phenol by following the pH change of an acid–base titration. Comparison of this curve with a curve obtained from titration of a synthetic mixture containing an amount of phenol corresponding to about 1% of the initial concentration of phenylacetic acid used in the experiment showed that negligible, if any, amounts of phenol were produced in the decomposition.

TABLE II

Effect of packed reaction vessel on % decomposition of phenylacetic acid, % CO, k, and k_1

Expt. No.	T, °A	Total % decomp.	% CO	k (sec^{-1})	k_1 (sec^{-1})
119	876	11.4	8.8	.35	.13
136P	876	71.5	52.5	2.98	.18
97	881	30.4	11.1	.43	.17
131P	881.5	57.0	45.0	2.45	.22
116	892	17.7	6.2	.50	.23
135P	892	77.6	1.4	3.88	.31
94	920	37.9	9.2	1.58	.70
134P	920	87.1	50.5	6.18	.78
103	939.5	34.7	2.7	1.22	1.00
133P	940	90.5	48.4	7.83	1.42

Attempts were made to detect formic acid in the reaction products but the large excess of toluene and the presence of other reaction products made the analysis extremely difficult and positive identification was not made.

From two experiments in the packed vessel the contents of the trap maintained at −5° C were titrated by following the pH change, and the presence of a small amount of a strong acid was evident. Although positive identification was not made, this acid was suspected to be oxalic. No evidence for the formation of oxalic acid was observed in the unpacked vessel.

DISCUSSION

Mechanism of Decomposition

The results suggested that the decomposition might occur by two simultaneous processes,

$$C_6H_5CH_2COOH \xrightarrow{k_1} C_6H_5CH_2. + .COOH, \qquad [1]$$

$$C_6H_5CH_2COOH \xrightarrow{k_2} C_6H_5CHCO + H_2O, \qquad [2]$$

reaction [1] being predominant. If each carboxyl radical is assumed to yield subsequently one molecule of either carbon dioxide or carbon monoxide, the rate of production of (CO_2 + CO) would be a measure of the rate of reaction [1]. To calculate the rate constant of this step, it was assumed that both processes [1] and [2] were first-order with respect to the acid concentration. The rate of decomposition of acid, A, may then be written

$$-\frac{dA}{dt} = (k_1+k_2)A = kA$$

where k is the over-all rate constant for the disappearance of acid. The rate of production of carbon dioxide plus carbon monoxide is then given by

$$\frac{d(CO_2+CO)}{dt} = k_1A = k_1(A_0e^{-kt})$$

and on integration it follows that

$$[CO_2+CO] = \frac{k_1}{k}A_0(1-e^{-kt}) = \frac{k_1}{k}(A_0-A).$$

The rate constant k was calculated in terms of the amount of acid decomposed and rate constant k_1 could then be evaluated from the measured yield of (CO_2 + CO). It ought to be pointed out that this calculation is not sensitive to the order of reaction [2], especially at low conversions, and also does not depend on an analysis for phenylketene. The rate constant k_2 represents the difference between k and k_1. Because of the large error necessarily involved in determining a small difference between two rate constants, no attempt was made to calculate an activation energy for reaction [2].

The first-order rate constants, k and k_1, measured for a variety of experimental conditions are listed in Table I. The rate constant k_1 remained almost unaffected, within the experimental accuracy, when the reaction time was varied by a factor of 4, the partial pressure of acid by a factor of 3, and the total pressure of toluene by a factor of 3. These results are shown in Tables IV, V, and VI. The small increase of about 30% in k_1

TABLE III

Effect of variation of toluene pressure on composition of non-condensable gases

Expt. No.	T, °A	Toluene pressure, mm	% CH_4	% CO	% H_2
98	879	6.3	28.8	13.3	57.9
96	881	12.1	30.6	11.2	58.2
112	956	4.7	20.2	10.1	69.8
113	957	11.1	24.0	6.6	69.4
106	968	4.9	18.3	7.4	74.4
107	968.5	10.0	24.1	6.1	69.8
92	990	6.6	15.1	19.1	65.8
91	991	10.9	20.3	17.3	62.3

in the packed reaction vessel is not considered significant because the surface-to-volume ratio had been increased by a factor of about 27. Therefore, the rate of reaction [1], as measured by the rate of production of carbon dioxide plus carbon monoxide, is considered to be a homogeneous, first-order dissociation process. The plot of log k_1 vs. $1/T$ was represented by a straight line over the range of temperature studied and is shown in Fig. 1. From this line the activation energy of reaction [1] was calculated as 55 kcal/mole and the frequency factor as 8×10^{12} sec^{-1}.

The largest source of error in the measurements was considered to be the uncertainty in the temperature of the reaction vessel during an experiment, due both to fluctuations in the temperature of the furnace as a whole and to variation in temperature along the length of the reaction vessel. The error in the rate constant k_1 resulting from these fluctuations was estimated to be about 8%. This would give an uncertainty of not more than 1 kcal/mole to the activation energy for reaction [1] calculated from the slope of the plot of log k_1 against $1/T$. However, in kinetic studies such as this there is an uncertainty, inherent in the assumption of a reaction mechanism, which prevents the unequivocal identification of the measured rate with the rate of the postulated processes. Part of this uncertainty arises from side reactions which may affect the products by which the rate of the primary step is measured. Thus, an activation energy assigned to a specific reaction probably contains errors larger than those estimated solely from the measurements involved. Unfortunately, the magnitude of such errors is very difficult to assess.

TABLE IV

Effect of reaction time on first-order rate constant k_1

Expt. No.	T, °A	Time of contact, sec	k_1 (sec^{-1})
120	881	.36	.15
98	879	.37	.17
96	881	.49	.16
97	881	.85	.17
125	882.5	.95	.15
111	904	.29	.40
109	904	.31	.36
110	904	.41	.36
126	904	.975	.34
94	926	.30	.69
95	926	.47	.74
127	928	1.05	.65

TABLE V

Effect of partial pressure of acid on first-order rate constant k_1

Expt. No.	T, °A	Acid pressure, mm Hg	k_1 (sec^{-1})
98	879	0.10	0.17
97	881	0.18	0.17
96	881	0.21	0.16
120	881	0.29	0.15
100	938	0.11	.95
101	938.5	0.20	1.05
102	939	0.34	1.03
83	947	0.12	1.47
84	950	0.38	1.64

TABLE VI

Effect of toluene pressure on first-order rate constant k_1

Expt. No.	T, °A	Toluene pressure, mm Hg	k_1 (sec^{-1})
96	881	12.1	.16
98	879	6.3	.17
116	892	12.7	.23
115	892	4.7	.22
94	926	12.9	.69
95	926	5.3	.74
113	957	11.1	1.81
112	956	4.7	1.80
107	968.5	10.9	2.64
106	968	4.9	2.60
91	991	11.0	5.61
92	990	6.6	5.52

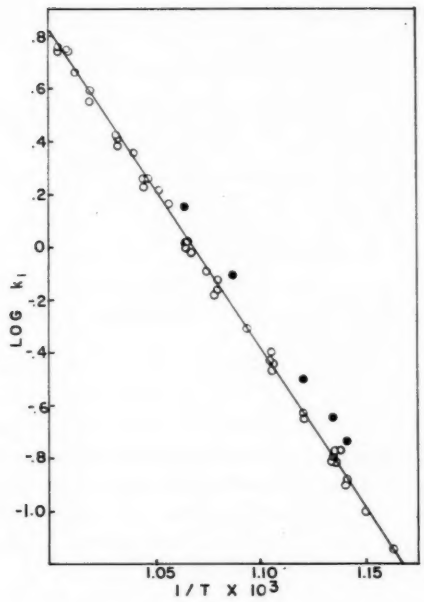

FIG. 1. Plot of log k_1 against $1/T$ for the thermal decomposition of phenylacetic acid.

○Experiments done in the unpacked reaction vessel.

●Experiments done in the packed reaction vessel.

Fate of the Carboxyl Radical

The carboxyl radical may decompose thermally in the following ways:

$$COOH \rightarrow CO_2 + H \quad [3]$$

$$COOH \rightarrow CO + OH \quad [4]$$

According to this scheme each carbon dioxide molecule is accompanied by a hydrogen atom, and each carbon monoxide molecule by a hydroxyl radical. On the basis of previous results (9) it was expected that the hydrogen atoms would react with toluene to give hydrogen and methane in the ratio 3:2.

$$H + C_6H_5CH_3 \rightarrow C_6H_5CH_2 + H_2 \quad [5a]$$

$$H + C_6H_5CH_3 \rightarrow C_6H_6 + CH_3 \quad [5b]$$

$$H + C_6H_5CH_3 \rightarrow C_6H_5 + CH_4 \quad [5c]$$

The most probable reaction of hydroxyl radicals with toluene was thought to be abstraction of a hydrogen atom to form water.

$$OH + C_6H_5CH_3 \rightarrow C_6H_5CH_2 + H_2O \quad [6a]$$

However, the displacement reactions, similar to reactions [5*b*] and [5*c*],

$$OH + C_6H_5CH_3 \rightarrow C_6H_5OH + CH_3 \quad [6b]$$

$$OH + C_6H_5CH_3 \rightarrow C_6H_5 + CH_3OH \quad [6c]$$

might also participate. Failure to detect phenol showed that reaction [6*b*] did not occur.

It was therefore assumed that hydroxyl radicals reacted with toluene to give products trapped at −78° C.

If all the benzyl radicals dimerized to give dibenzyl,

$$2C_6H_5CH_2 \rightarrow (C_6H_5CH_2)_2, \qquad [7]$$

the following stoichiometry should have been observed:

$$CO_2 = H_2 + CH_4$$
$$(C_6H_5CH_2)_2 = CO_2 + CO = H_2 + CH_4 + CO$$
$$H_2/CH_4 = 1.5$$

None of these conditions was in fact observed. As mentioned previously, the carbon dioxide was always greater than the sum of the hydrogen and methane, and the ratio of hydrogen to methane varied from 2.0 to 3.5. The first-order rate constant calculated from the amount of ($H_2 + CH_4 + CO$) produced was always lower than that calculated from the yield of ($CO_2 + CO$) and varied with reaction time. Moreover, the Arrhenius plot of the former rate constant was not rectilinear over the whole temperature range studied. The yield of dibenzyl was always much less than that of the carbon dioxide plus carbon monoxide, but was usually between 75 and 90% of the non-condensable gases (Table I).

Reactions [3] to [6] are therefore insufficient to account for the products formed. If carboxyl radicals do not all decompose rapidly by reactions [3] and [4], they may react with each other according to reactions [8] or [9].

$$2COOH \rightarrow 2CO_2 + H_2 \qquad [8]$$
$$2COOH \rightarrow CO_2 + CO + H_2O \qquad [9]$$

Either of these reactions produces one molecule of carbon dioxide or carbon monoxide for each carboxyl radical, and the formation of ($CO_2 + CO$) is still a measure of the extent of reaction [1]. However, only one molecule of non-condensable gas is produced for every two carboxyl radicals, and the total amount of non-condensable gases will be less than the sum of the carbon dioxide and carbon monoxide. Reaction [8] also produces molecular hydrogen unaccompanied by methane and will result in a ratio of hydrogen to methane greater than that observed from the reaction of hydrogen atoms with toluene. The indication that under certain conditions oxalic acid may have been a product of the decomposition in the packed reaction vessel suggests that carboxyl radicals have a sufficiently long lifetime under the conditions of the present experiments to react with each other.

If carboxyl radicals react with each other, then reaction between carboxyl and benzyl radicals may be considered at least as probable. If the radicals recombined to reform phenylacetic acid, the rate constant k_1 would not appear to be first-order over the temperature range studied and under the different conditions of decomposition. Therefore, the recombination between benzyl and carboxyl radicals is not considered significant. An alternative and possibly more probable reaction for benzyl and carboxyl radicals would be the disproportionation reaction

$$C_6H_5CH_2 + COOH \rightarrow C_6H_5CH_3 + CO_2, \qquad [10]$$

which is exothermic to the extent of about 63 kcal/mole. This value represents the difference between $D(C_6H_5CH_2—H)$ = 83 kcal/mole (11) and $D(COO—H)$ = 20 kcal/mole. The latter value is deduced in the last paragraph of this paper.

Reactions [8], [9], and [10] lead to an excess of carbon dioxide relative to dibenzyl or

hydrogen plus methane, but the total amount of non-condensable gases should equal the amount of dibenzyl.

$$CO_2 + CO > (C_6H_5CH_2)_2 = H_2 + CH_4 + CO$$

The latter condition was indeed found to hold within the experimental accuracy for the recovery of dibenzyl. The ratio of dibenzyl to non-condensable gases varied between 0.75 and unity (Table I). In previous studies it was shown that dibenzyl was not trapped quantitatively at about $-5°$ C from a stream of toluene, and the recoveries were of the order of 80–90% of the expected amount (10). It was considered, therefore, that the yield of dibenzyl was probably equal to the yield of non-condensable gases ($H_2 + CH_4 + CO$).

Reaction of the carboxyl radical with toluene may also occur, probably to produce carbon dioxide and hydrogen, or carbon monoxide and water.

$$COOH + C_6H_5CH_3 \rightarrow C_6H_5CH_2 + (HCOOH) \begin{cases} \nearrow CO_2 + H_2 & [11a] \\ \searrow CO + H_2O & [11b] \end{cases}$$

These reactions yield one molecule of carbon dioxide or carbon monoxide for each carboxyl radical, but they also produce a molecule of dibenzyl for each carbon dioxide or carbon monoxide formed. If formic acid were the product of the reaction, the yield of dibenzyl relative to that of ($CO_2 + CO$) would be even greater than if carbon dioxide or carbon monoxide were formed only by reactions [3] and [4].

Moreover, if reaction [11*a*] or [11*b*] occurred, each would be favored by an increase in toluene pressure and more hydrogen or carbon monoxide would be expected to be produced at the expense of methane under these conditions. However, as shown in Table III, this trend was not observed.

In previous studies with the toluene-carrier technique the production of dibenzyl was taken as a measure of the extent of the decomposition of the reactant into free radicals. The small amounts of dibenzyl formed relative to carbon dioxide may be interpreted, therefore, as an indication of an alternate mode of decomposition of phenylacetic acid by a molecular rearrangement into carbon dioxide and toluene.

$$C_6H_5CH_2COOH \rightarrow CO_2 + C_6H_5CH_3 \qquad [12]$$

If reaction [12] occurred simultaneously with reaction [1] the plot of log k_1 vs. $1/T$ would be expected to be a composite curve representing the sum of the rate constants of two processes which would probably require different activation energies. Because the plot of log k_1 vs. $1/T$ was represented by a straight line it would appear that if the molecular decomposition occurred it must have the same activation energy and the same frequency factor as reaction [1]. Therefore, although the occurrence of reaction [12] cannot be ruled out unequivocally,* its participation would not affect the conclusions derived for reaction [1], which is represented by the homogeneous, unimolecular dissociation process,

$$C_6H_5CH_2COOH \rightarrow C_6H_5CH_2 + COOH.$$

The activation energy of this step was calculated as 55 kcal/mole and the frequency factor as 8×10^{12} sec^{-1}. With the usual assumption that the recombination of radicals requires no activation energy, the value of 55 kcal/mole may be identified with the dissociation energy of the $C_6H_5CH_2$—COOH bond.

**It is possible to visualize that if reaction [12] occurred to a small extent, the linearity of the Arrhenius plot may not be seriously affected though the activation energy calculated from it may be erroneous.*

The heat of formation of carboxyl radical can be calculated using the thermochemical relation

$$\Delta H_f(\text{COOH}) = \Delta H_f(\text{C}_6\text{H}_5\text{CH}_2\text{COOH}) + D(\text{C}_6\text{H}_5\text{CH}_2\text{—COOH}) - \Delta H_f(\text{C}_6\text{H}_5\text{CH}_2).$$

Although the value for the heat of formation of benzyl radical is somewhat uncertain, a reasonable estimate for it seems to be 42 kcal/mole (11). No recent determination of the heat of formation of phenylacetic acid in its monomeric form in the gaseous phase is available. However, a value of -75 kcal/mole for $\Delta H_f(\text{C}_6\text{H}_5\text{CH}_2\text{COOH})$ may be computed from relevant thermochemical data.* In this calculation, the present authors have assumed that the heat of vaporization determined by Stull (12) corresponds to the heat of vaporization of phenylacetic acid into monomers. Taking for $D(\text{C}_6\text{H}_5\text{CH}_2\text{—COOH})$ the value of 55 kcal/mole derived in this study, it follows that $\Delta H_f(\text{COOH}) \simeq -62$ kcal/mole. In turn this value leads to the following dissociation energies $D(\text{COO—H}) = 20$ kcal/mole and $D(\text{CO—OH}) = 45$ kcal/mole.

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*$\Delta H_{\text{combustion}}(\text{C}_6\text{H}_5\text{CH}_2\text{COOH})_s = 930.2$ *kcal/mole (13)*, $\Delta H_{\text{fusion}}(\text{C}_6\text{H}_5\text{CH}_2\text{COOH}) = 30$ *cal/g (14)*, $\Delta H_{\text{vaporization}}(\text{C}_6\text{H}_5\text{CH}_2\text{COOH}) = 16.4$ *kcal/mole (12)*, $\Delta H_f(\text{CO}_2)_g = -94.05$ *kcal/mole (15)*, $\Delta H_f(\text{H}_2\text{O})_l = -68.32$ *(15). From these data it follows that the heats of formation of phenylacetic acid in the solid, liquid, and gaseous forms are* -95.48, -91.4, *and* -75 *kcal/mole, respectively. For the calculation of* $D(\text{CO}_2\text{—H})$ *and* $D(\text{CO—OH})$ *the following heats of formation were used:* $\Delta H_f(\text{CO})_g = -26.4$ *kcal/mole (15)*, $\Delta H_f(\text{H}) = 52.1$ *kcal/mole (15), and* $\Delta H_f(\text{OH}) = 9.5$ *kcal/mole based on the value of 101.3 kcal/mole (16) for* $D(\text{O—H})$.

THE THERMAL DECOMPOSITION OF DIPHENYLACETIC ACID[1]

MARGARET H. BACK[2] AND A. H. SEHON

ABSTRACT

The pyrolysis of diphenylacetic acid was investigated over the temperature range 515–636° C using the toluene-carrier technique. The main products of this system were similar to those formed in the thermal decomposition of phenylacetic acid. The rate constant for the reaction

$$(C_6H_5)_2CH.COOH \xrightarrow{k_1} (C_6H_5)_2CH + COOH$$

was calculated as $k_1 = 8\times10^{12}.\exp(-52{,}000/RT)$ sec^{-1}.

The over-all decomposition reaction was found to be partly heterogeneous and, therefore, the activation energy for reaction [1] may be only tentatively identified with $D[(C_6H_5)_2CH—COOH]$.

The investigation described in the foregoing paper (1) was extended to the thermal decomposition of diphenylacetic acid in a stream of toluene for the purpose of producing carboxyl radicals from another source and for studying their reactions further. On theoretical grounds the diphenylmethyl radical ought to be stabilized by resonance to a larger extent than the benzyl radical. Therefore, it was expected that the bond $(C_6H_5)_2CH—COOH$ would be weaker than the bond $C_6H_5CH_2—COOH$, and that diphenylacetic acid would decompose more readily than phenylacetic acid.

EXPERIMENTAL

The apparatus and procedure used were identical with those employed for the thermal decomposition of phenylacetic acid (1).

Diphenylacetic acid (m.p. 140° C) was obtained from Eastman Kodak Co. and used without further purification. The toluene was generously supplied by Gulf Oil Corporation, Pittsburgh, Pa. It was purified by treatment with dilute sodium hydroxide for 24 hours, dried, and distilled through an efficient column. The fraction boiling at 110.5° C was used in this study.

RESULTS

The products from the decomposition of diphenylacetic acid were similar to those from phenylacetic acid and are shown in Table I. Carbon dioxide was again the major product and carbon monoxide a minor product only. The sum of the hydrogen and methane produced amounted to 20–50% of the carbon dioxide, and the ratio of hydrogen to methane varied from 1.9 to 4.5, both trends increasing with increasing temperature. Carbon monoxide accounted for 30–40% of the non-condensable gases, compared with about 10% in the pyrolysis of phenylacetic acid.

The sum of the carbon dioxide and carbon monoxide produced and the undecomposed acid recovered accounted for 94–98% of the initial amount of acid. Therefore, analysis was not attempted, as with phenylacetic acid (1), for other possible products. However, small amounts of a brilliant yellow compound, suspected to be diphenylketene, were observed in the trap maintained at −5° C. The contents of this trap were dissolved in alcohol, since the latter is known to react rapidly with diphenylketene to form the ester

[1]*Manuscript received April 11, 1960.*
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TABLE I

Products and rate constants of the decomposition of diphenylacetic acid

Expt. No.	T, °A	t, sec	Total pressure, mm Hg	Pressure acid, mm Hg	% decomp.	CO_2, mmoles	Non-condensable gases, mmoles	% CH_4	% CO	% H_2	H_2/CH_4	k, sec^{-1}	k_1, sec^{-1}
3P	788	.57	11.1	0.09	30.2	.037	.011	4.9	82.4	17.6	3.6	.63	.08
14	804	.695	5.55	0.1	9.3	.044	.009	13.5	51.6	34.9	2.6	.14	.06
16	805	.56	11.5	0.07	9.0	.040	.010	13.3	42.8	43.9	3.3	.17	.06
39	806	.87	10.6	0.12	8.9	.068	.012	21.2	38.3	40.5	1.9	.11	.07
35	807	1.75	10.8	0.18		.128	.029	14.0	47.4	38.6	2.8		.08
38	813	.82	10.2	0.13	11.6	.087	.015	20.5	36.9	42.5	2.1	.15	.10
29	821	.48	10.5	0.14	11.0	.057	.013	17.5	35.0	47.5	2.7	.24	.12
42	824.5	.87	10.2	0.30	15.4	.133	.027	18.4	30.3	51.3	2.8	.19	.13
41	826	.89	10.3	0.12	16.0	.125	.028	22.1	29.4	48.5	2.2	.20	.14
28	835	.46	4.05	0.13	15.6	.0885	.023	16.9	34.9	48.2	2.85	.37	.21
40	836.5	.87	10.4	0.11	22.0	.183	.044	21.2	30.3	48.5	2.3	.28	.25
2P	844	.47	10.5	0.08	61.1	.167	.0785	5.3	81.0	13.6	2.6	2.0	.71
30	849.5	.72	5.4	0.11	26.3	.192	.063	15.6	32.7	51.6	3.3	.42	.33
18	850	.49	11.9	0.15	23.0	.172	.051					.53	.37
17	851	.475	11.9	0.06	23.8	.142	.045	19.5	34.8	46.2	2.4	.57	.40
5P	852	.67	5.5	0.14	73.2	.207	.142	3.2	85.2	11.6	3.6	2.0	.81
31	866	.44	11.0	0.19	28.8	.219	.074	18.0	27.2	54.8	3.0	.77	.61
4P	867.5	.45	10.3	0.08	84.4	.203	.181	4.3	83.8	11.9	2.8	4.17	1.99
25	868.5	.555	5.1	0.08	35.1	.240	.099	15.6	33.6	50.8	3.3	.78	
26	869	.66	5.7	0.08	39.3	.334	.140	14.6	32.2	53.2	3.6	.76	.68
27	870.5	.51	11.7	0.10	35.4	.304	.108	18.3	29.4	52.4	2.9	.85	.81
32	877	.72	5.3	0.09	50.5	.409	.183	12.4	33.0	54.6	4.4	.98	.90
33	877	.51	12.9	0.15	40.0	.344	.131	17.4	23.2	59.4	3.4	1.00	.90
19	880.5	.45	11.15	0.05	44.2	.351	.163	19.7	28.8	51.4	2.6	1.29	1.10
20	880.5	.46	10.4	0.18	42.6	.366	.133	14.1	28.6	57.3	4.1	1.21	1.10
1P	897	.41	10.6	0.07	86.6	.324	.433	5.6	78.1	16.4	2.9	4.87	3.87
24	897	.43	10.7	0.14	55.7	.496	.225	14.6	27.5	57.9	4.0	1.90	1.80
23	897.5	.42	11.1	0.13	57.5	.619	.235	14.7	25.2	60.2	4.1	2.04	
21	903	.43	10.7	0.14	66.6	.610	.287	13.1	27.6	59.4	4.5	2.57	2.50
34	904	.42	10.7	0.13	62.6	.568	.278	14.3	24.3	61.4	4.3	2.32	2.19
22	909	.43	10.95	0.13	71.2	.629	.312	13.2	26.8	60.0	4.55	2.93	2.82

(2). In consequence, the ketene was not expected to interfere with the titration for the undecomposed acid. The weight of diphenylketene was obtained by assuming that this compound accounted for the difference between the acid decomposed and the (CO_2 + CO) recovered.

The recombination of benzyl and diphenylmethyl radicals may produce three possible products: dibenzyl, 1,1,2,2-tetraphenylethane, and 1,1,2-triphenylethane. A white solid, insoluble in alcohol, was deposited in the trap at −5° C, and it is suggested that this might be 1,1,2,2-tetraphenylethane, since the other recombination products are soluble in alcohol. The weight of the recombination products was estimated by subtracting the weights of undecomposed acid and of diphenylketene from the total weight of products trapped at −5° C. The moles of recombination products were obtained by assuming that benzyl and diphenylmethyl radicals were formed in equal amounts. Although the errors were necessarily large, the quantities obtained were consistently lower than the (CO_2 + CO) but slightly higher than the total amount of non-condensable gases. This will be discussed later.

No acid was found in the trap kept at −78° C, and it was concluded that formic acid was not a product of the reaction. This lends support to the conclusion that formic acid was also not a product of the decomposition of phenylacetic acid, which was previously deduced by indirect evidence in the thermal decomposition of phenylacetic acid.

DISCUSSION

These results suggest that the two modes of decomposition postulated for phenylacetic acid also occurred in the decomposition of diphenylacetic acid, but that in this case reaction [2] was of much less importance.

$$(C_6H_5)_2CHCOOH \xrightarrow{k_1} (C_6H_5)_2CH + COOH$$

$$(C_6H_5)_2CHCOOH \xrightarrow{k_2} (C_6H_5)_2C{=}C{=}O + H_2O$$

Since the general pattern of the decomposition of diphenylacetic acid was essentially similar to that of phenylacetic acid, it appeared that carboxyl radicals formed from the decomposition of diphenylacetic acid disappeared by the same reactions postulated in the decomposition of phenylacetic acid (1). Decomposition of the carboxyl radical into carbon dioxide is energetically more favorable than decomposition into carbon monoxide, and it is unlikely that the activation energy for the homogeneous reaction to give carbon dioxide would be greater than for the reaction producing carbon monoxide. However, carbon monoxide formed a larger fraction of the non-condensable gases from diphenylacetic acid than from phenylacetic acid. Furthermore, the ratio of carbon monoxide to carbon dioxide, although it was always less than one, was greater at the lower temperatures used for the decomposition of diphenylacetic acid. It would seem, therefore, that at least one of the reactions giving carbon monoxide has a lower activation energy than those giving carbon dioxide. With both acids the production of carbon monoxide was increased in the packed vessel (Table II) and was also favored at low toluene pressures

TABLE II

Effect of packed reaction vessel on % decomposition of diphenylacetic acid, % CO, k,* and k_1

Expt. No.	T, °A	Total % decomp.	% CO	k, sec^{-1}	k_1, sec^{-1}
30	849.5	26.3	32.7	.42	.42
2P	844	61.1	81.0	2.04	1.33
17	851	23.8	34.8	.57	.40
5P	852	73.2	85.2	1.97	1.15
31	866	28.8	27.2	.77	.61
4P	867.5	84.4	83.8	4.17	1.86
24	897	55.7	27.5	1.90	1.80
1P	897	86.6	78.1	4.87	2.86

*k refers to the first-order rate constant for the over-all reaction calculated in terms of the acid decomposed.

(Table III). Consequently, it is suggested that carboxyl radicals decomposed primarily heterogeneously to yield carbon monoxide and a hydroxyl radical. The activation energy for such a reaction might be fairly low.

The interaction of carboxyl radicals by reactions [8] and [9] of the scheme postulated in the previous paper (1) leaves an excess of diphenylmethyl radicals over benzyl radicals. Thus the weight of product condensed at 0° C would be greater than it would be if diphenylmethyl radicals and benzyl radicals were formed in equal molar ratios. This would make the estimation of the moles of recombination product, based on a 1:1 molar ratio of diphenylmethyl and benzyl radicals, too high. This might explain why the calculated

TABLE III

Effect of variation of toluene pressure on composition of non-condensable gases of diphenylacetic acid

Expt. No.	T, °A	Toluene pressure, mm Hg	% CH_4	% CO	% H_2
32	877	5.3	12.4	33.0	54.6
33	877	12.9	17.4	23.2	59.4
25	868.5	5.1	15.6	33.6	50.8
31	866	11.0	18.0	27.2	54.8
14	804	5.6	13.5	51.6	34.9
39	806	10.6	21.2	38.3	40.5

amounts of recombination products were greater than the amounts of non-condensable gases, although the results are not sufficiently accurate to permit a quantitative estimate of the effect.

The rate constant k_1 was calculated from the rate of production of carbon dioxide plus carbon monoxide, assuming both reactions [1] and [2] to be first-order. The results are shown in Table I. This rate constant was practically unaffected by varying the partial pressure of acid by a factor of 3 (Table IV), the reaction time by a factor of 2

TABLE IV

Variation of k_1 with partial pressure of acid

Expt. No.	T, °A	Acid pressure, mm Hg	k_1, sec^{-1}
16	805	0.07	.08
39	806	0.12	.07
35	807	0.18	.08
41	826	0.12	.14
42	824.5	0.30	.13
17	851	0.06	.40
18	850	0.15	.37
19	880.5	0.04	1.10
20	880.5	0.17	1.10

(Table V), and the total pressure of toluene by a factor of 2 (Table VI). The plot of log k_1 against $1/T$ was represented by a straight line over the range of temperatures studied and

TABLE V

Variation of k_1 with reaction time

Expt. No.	T, °A	t, sec	k_1, sec^{-1}
16	805	.56	.06
39	806	.87	.07
29	821	.48	.12
41	826	.89	.14
28	835	.46	.21
40	836.5	.87	.23

TABLE VI

Variation of k_1 with toluene pressure

Expt. No.	T, °A	Toluene pressure, mm Hg	k_1, sec^{-1}
14	804	5.55	.06
16	805	11.5	.06
26	869	5.6	.68
27	870.5	11.7	.81
32	877	5.3	.90
33	877	12.9	.90

is shown in Fig. 1. From the slope of this line an activation energy of 52 kcal/mole and a frequency factor of 8×10^{12} sec^{-1} were calculated.

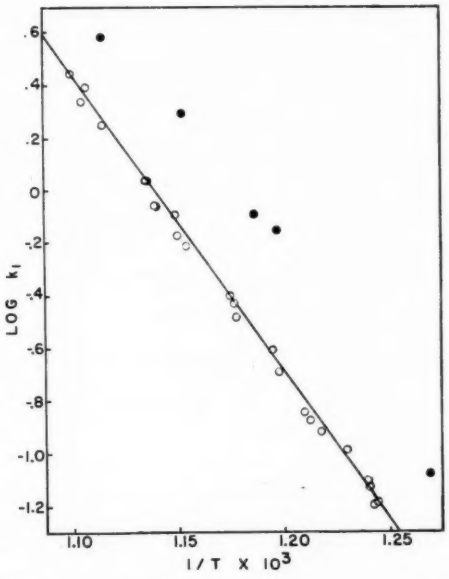

FIG. 1. Plot of log k_1 against $1/T$ for the thermal decomposition of diphenylacetic acid.

○ Experiments done in the unpacked reaction vessel.

● Experiments done in the packed reaction vessel.

The pyrolysis of diphenylacetic acid appeared to be more affected by the surface conditions than that of phenylacetic acid. In the packed reaction vessel the over-all decomposition was greatly increased and much larger quantities of the yellow compound, suspected to be diphenylketene, were observed. The rate constant calculated from the amount of (CO_2 + CO) produced was increased by a factor of almost three as compared with the values obtained with the unpacked reaction vessel (Table II). However, the slope of the plot of log k_1 against $1/T$ was similar to that obtained with the values calculated for the unpacked reaction vessel. Since in the packed reaction vessel the surface to volume ratio was increased by a factor of about 27, an increase in k_1 by a factor of 3 may be interpreted as an indication that only about 11% of the free radical decomposition occurred on the surface of the vessel. Therefore, the pyrolysis of diphenylacetic acid may be considered to be primarily a homogeneous reaction.

Because of the uncertainty of the role of the surface in this decomposition, the identification of the activation energy with the bond dissociation energy, $D[(C_6H_5)_2CH—COOH]$, is somewhat tentative. However, the frequency factor is close to that expected for a unimolecular dissociation and, remarkably, it is identical with that for the decomposition of phenylacetic acid. Moreover, the value of 52 kcal/mole is reasonable considering that the bond dissociation energy assigned to the corresponding bond in phenylacetic acid was estimated at 55 kcal/mole. The difference between these two bond dissociation

energies is supported by the results obtained from the studies of the thermal decompositions of benzyl and diphenylmethyl bromides (3). In the latter studies the diphenylmethyl radical was found to be slightly more stabilized by resonance than the benzyl radical.

ACKNOWLEDGMENTS

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CLAM POISON

I. THE PAPER CHROMATOGRAPHIC PURIFICATION OF CLAM POISON DIHYDROCHLORIDE[1]

A. A. CASSELMAN, R. GREENHALGH, H. H. BROWNELL, AND R. A. B. BANNARD

ABSTRACT

A semimicro (10–25 mg) paper chromatographic method for purification of clam poison has been developed. This procedure when applied to a partially purified (toxicity 4150 MU/mg) sample of the toxin gave a 74% recovery of clam poison dihydrochloride of purity identical with that of the best material obtained by Schantz *et al.* (2) via the alumina chromatographic method. The upper limit of purity of toxin resulting from paper chromatography is estimated to be 98% because of traces of impurity introduced by the paper. The paper chromatographic procedure has been found capable of enriching low-grade toxic residues (35%) from the alumina process to a purity of 75% with an over-all recovery of 70%.

INTRODUCTION

Mold and co-workers (1) recently described the criteria for purity of clam poison dihydrochloride, the paralytic poison which can be isolated from the siphon of the toxic Alaska butter clam *Saxidomus giganteus* (2). Among other properties, the paper chromatography of clam poison dihydrochloride was examined in several solvent systems and it was observed (1) that the compound gave only one typical pink spot with Weber reagent (3) (R_F 0.30) when phenol–water (4:1) or *t*-butanol – acetic acid – water (2:1:1) was used as developer. The above-mentioned workers kindly provided us with 300 mg of partially purified (toxicity 4150 MU*/mg) clam poison which we wished to use for structural investigation. Paper chromatograms revealed, however, that the sample contained at least five components and it became evident to us that further purification was necessary. Because of the small quantity at our disposal, the method of purification used by Schantz *et al.* (2), involving as its final step chromatography on acid-washed alumina, was considered likely to be too wasteful due to the relatively low recovery of material of high toxicity (25–45%). Furthermore, prior to commencing degradative studies it was considered necessary to develop a semimicro method applicable to reaction mixtures from clam poison which would make feasible both the isolation of products and recovery in a high state of purity of any unreacted toxin. In addition, it was desired to examine the feasibility of paper chromatography as a means for purification of the clam poison residues of low toxicity (hereinafter designated as clam poison tailings) which on reworking by a variety of chromatographic methods (4) had failed to produce material with an enhanced toxicity. The work now to be outlined describes the development of the paper chromatographic method of purification and its application to (*a*) recovery and further investigation of the components contained in the partially purified sample of toxin furnished by Schantz *et al.* (2), and (*b*) enrichment of clam poison tailings.

RESULTS AND DISCUSSION

Chromatography of the partially purified clam poison dihydrochloride sample on *unwashed* Whatman No. 1 paper, using *t*-butanol – acetic acid – water (2:1:1)† as

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*MU = *mouse unit.*
† *All chromatograms were developed in this solvent system unless otherwise specified.*

developer gave three pink spots, R_F 0.19, 0.32, and 0.53, on spraying with Weber (3) reagent. These components could also be detected as blue spots by spraying with bromcresol green, or as fluorescent spots on exposure to ultraviolet light. Ultraviolet light also revealed the presence of another component, R_F 0.0. When the chromatograms were sprayed with Sakaguchi (5, 6) reagent, which is specific for monoalkylguanidines (7), a typical red spot appeared at R_F 0.70, indicating the presence of a fifth component. Component 3 (see Table I) appeared as an atypical greenish yellow spot with Sakaguchi reagent and as an atypical yellow spot with ninhydrin. Resolution of the clam poison sample into five components was also observed when other solvent systems were used, as shown in Table I. From these results, and those of Mold *et al.* (1), it seemed probable that component 3 is clam poison and that the other components present in the sample are impurities either present initially or formed by partial decomposition during storage in aqueous solution at pH 5.

TABLE I

Paper chromatography of clam poison dihydrochloride

	R_F in solvent system				
Component	*t*-Butanol – acetic acid – water (2:1:1)	Phenol–water (4:1)	*t*-Butanol – 1% hydrochloric acid (3:1)	*n*-butanol – acetic acid – water (25:6:25)	*n*-Butanol – 1% hydrochloric acid (5:6)
1*	0.00	0.03	0.00	0.02	0.00
2*†‡	0.19	0.20	0.05	0.07	0.05
3*†‡	0.32	0.30	0.17	0.23	0.08
4*†‡	0.53	0.40	0.42	0.39	0.24
5§	0.70	0.91	0.64	0.55	0.24

*UV fluorescent. †Weber positive. ‡Bromcresol green positive. §Sakaguchi positive.

Application of paper chromatographic techniques to purification and recovery of the components present in the sample of partially purified clam poison dihydrochloride revealed that impurities were being introduced during the process, since greater than theoretical recoveries were at first obtained. The most probable source of such error is the presence of extractable impurities in the paper, or debris produced during chromatography by partial degradation of cellulose by the solvent system. Other workers (8, 9, 10, 11) have found that extensive purification or pretreatment of the paper is essential if reproducible results in quantitative paper chromatography are to be obtained where acidic or basic solvent systems are employed. Accordingly, several procedures for pretreatment of Whatman No. 1 paper were examined with the object of eliminating or minimizing extractable materials. After pretreatment according to the procedures given in Table II, blank runs were performed using only the pretreated paper and the developing solvent. After drying, strips corresponding in location and area to the five components previously found to be present in the sample were cut from the "chromatograms", eluted with water, and the eluates concentrated, lyophilized, and weighed. The amount of residue thus obtained should correspond to the amount of impurity introduced into a given component during actual purification. The results presented in Table III make it evident that procedure 2 (see Table II) yields eluates containing the smallest amounts of impurities, and this method was adopted for all subsequent quantitative work.

In Table III, three sets of results are given under procedure 2. The first column represents minimum quantities of impurities introduced by the method. The next two

TABLE II

Washing procedures for Whatman No. 1 paper

Procedure 1		Procedure 2		Procedure 3		Procedure 4	
Solution	Duration (days)	Solution	Duration (days)	Solution	Duration (days)	Solution	Duration (days)
0.01 *N* HCl	3	1 *N* Na_2CO_3	4.5	2 *N* HOAc	2.5	0.01 *N* HCl	3
H_2O	4	H_2O	2.5	H_2O	2	H_2O	4
		2 *N* HOAc	2	0.5 *N* LiOH	9	Phenol–water (4:1)	4
		H_2O	2	H_2O	5	H_2O	4
		0.1 *N* $Ca(OAc)_2$	2	1.5% 8-hydroxyquinoline (in 95% EtOH)	2.5		
		H_2O	2	95% EtOH	2.5		
		1.5% 8-hydroxyquinoline (in 95% EtOH)	2.5	Solvent system*	4		
		95% EtOH	4	H_2O	4		
		Solvent system*	4				
		H_2O	4				

*t-BuOH – HOAc – H_2O (2:1:1).

TABLE III

Impurity introduced during chromatographic purification of toxin components

		Weight of impurity (mg) for pretreatment					
Component No.	R_F	Procedure 1*	2*			3*	4†
1	0.00	0.24	0.17	0.19	0.22	0.46	0.31
2	0.19	0.29	0.14	0.25	0.21	0.40	0.36
3	0.32	0.59	0.19	0.26	0.23	0.37	0.49
4	0.53	0.56	0.25	0.51	0.50	0.52	0.37
5	0.70	1.05	0.26	0.41	0.37	0.81	0.37

*Solvent system *t*-butanol – acetic acid – water (2:1:1). †Solvent system phenol–water (4:1).

columns represent duplicate assays on another batch of paper pretreated in the same manner and are the maximum amounts of impurities encountered by the method. Thus, the quantity of impurity introduced per component is consistent for sheets within a given batch of paper, but some variation is to be expected from batch to batch. For actual purification of clam poison, therefore, blank runs were performed on two sheets from the batch of pretreated paper to be used in the purification. The arithmetic mean of the values thus obtained in the two runs for a given component was taken as the correction factor to be subtracted from the weight of that component recovered in actual chromatograms.

In procedure 2 the sodium carbonate treatment is used to open up the fiber structure of the paper and to remove hemicellulose (12), polypeptides (9), and acidic impurities. Treatment with acetic acid removes basic impurities and the alkaline earth metal ions Ca^{++} and Mg^{++} which can lead to ghost spots (8). Bad streaking effects have been observed, however, in the chromatography of amino acids and peptides (11) and the calcium acetate wash is used to overcome this disadvantage by replacement of some of the calcium ion removed in the preceding step. In acidic solvent systems, traces of heavy metal ions have been found responsible for the appearance of diffuse spots presumably due to formation of chelated complexes (8) and treatment with 8-hydroxyquinoline is used to remove these impurities. Finally, prolonged elution with the developing solvent

system is intended to minimize further degradation of the paper by the solvent during subsequent chromatography.

Further examination of the washing procedure revealed that no significant difference in quantities of extractables was caused by the use of the ascending rather than the descending technique. This result was somewhat unexpected since it was anticipated that the much longer development time required by the former method (40 hours vs. 16 hours) would lead to greater degradation of the paper and thus yield larger quantities of impurities. The fact that such did not occur attests to the effectiveness of the pretreatment process. Quantities of extractables were also not significantly affected by treating eluates with Amberlite IRA-400 (Cl) resin prior to lyophilization. The latter step is required during actual recovery of the toxin to ensure that it is completely converted to the dihydrochloride.

No difference in resolution of the components was observed by use of the two techniques referred to in the preceding paragraph and the descending method was adopted in all subsequent work because of the shorter development time required.

An effort was made by means of ultraviolet, fluorescence, and infrared spectral measurements to determine the nature of the impurities eluted from the pretreated paper in locations corresponding to components 1 to 5 in Table I. In the ultraviolet region, all the fractions showed weak absorption ($E^{1\%}_{1\,cm}$ = 8–28) at 255–260 mμ. Each fraction also exhibited a strong to medium intensity fluorescence at 350 mμ. The infrared spectra (KBr pellet) of all the fractions were very similar, but the bands (at 1000–1100 cm^{-1}; 1380–1410 cm^{-1}; 1570–1630 cm^{-1}; and 3100–3500 cm^{-1}) were broad rather than sharp. It was therefore not possible to draw any definite conclusion regarding the chemical structure of the impurities except that all the fractions, as expected, are of very similar if not identical composition.

For quantitative purification, 12–17 mg of partially purified clam poison dihydrochloride (toxicity 4150 MU/mg) was applied to specially pretreated Whatman No. 1 paper, chromatographed, and the components recovered and bio-assayed as described in the Experimental section. The results for a typical experiment, shown in Table IV, indicate that 96% of the material applied to the paper has been recovered, and that component 3, the only fraction which possesses significant toxicity, accounts for 74% of the sample. In other experiments performed in the same manner, 94–97% of the sample was accounted for with product distributions similar to that given in Table IV. The toxicity of component 3 by the mouse unit (MU) assay method (13) compares favorably with the value 5500±500 MU/mg reported by Schantz (2) and co-workers. In an LD_{50} assay, agreement was even better, the value found being 9.5 γ/kg (IP mice) vs. 10±1 γ/kg reported (14). In addition, the specific rotation ($[\alpha]^{25}_{D}$ +128°) was within the limits +130±5° given by Schantz *et al.* (2) and the substance gave an infrared spectrum identical with that reported (15, see Fig. 3). It is therefore evident that component 3 and the best material obtained by Schantz *et al.* by the alumina chromatographic method are of identical purity. The purity of component 3 was estimated on the basis of impurity known to be introduced from the paper to be 98% and in view of the supporting data given above this figure seems to be a realistic appraisal. Clam poison dihydrochloride, as recovered from the paper, was frequently obtained as a colorless to pale yellow glass. On two occasions, however, it was obtained as a mixture of glass and crystals with the latter predominating. The latter were very hygroscopic and did not melt but sintered above 200°.

Components 1, 2, 4, and 5 were not obtained in as high a degree of purity as clam poison because of the smaller quantities which were present and the relatively larger amount of

TABLE IV

Paper chromatographic purification of clam poison dihydrochloride

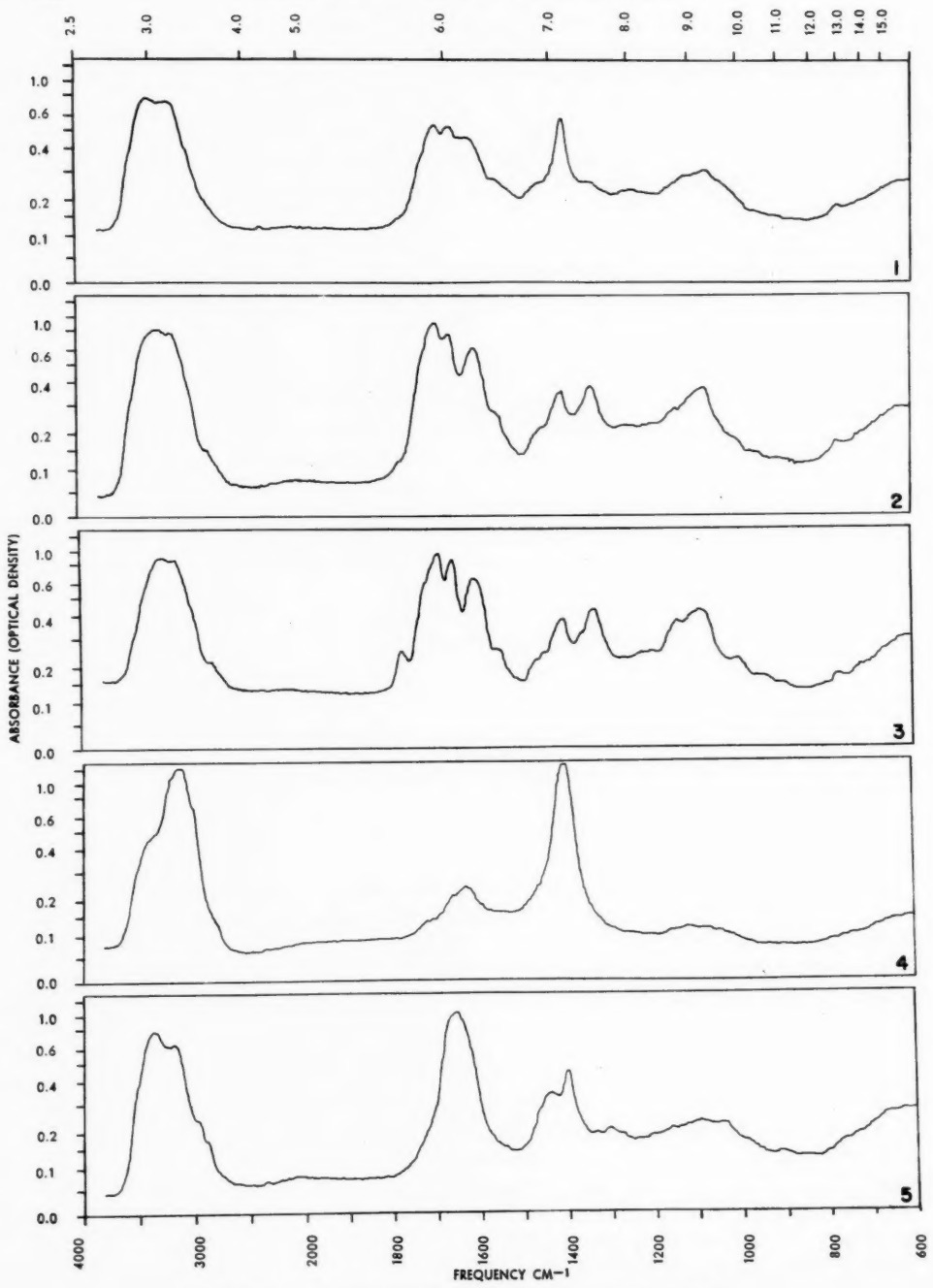
Component No.	R_F	Weight recovered (mg)	Corrected* weight (mg)	% recovery	Estimated % purity	Toxicity (MU/mg)	Dosage (mg/kg)
1	0.0	0.63	0.46	3.6	73	Nil	0.8
2	0.19	1.62	1.48	11.5	91	Nil	0.4
3	0.32	9.63	9.44	73.6	98	4800	0.018
4	0.53	0.84	0.59	4.6	70	Nil	1.0
5	0.70	0.65	0.40	3.1	61	Nil	1.4
			12.37†	96.4			

*Weight recovered — blank. †Theoretical recovery 12.83 mg.

impurity consequently introduced into each fraction by the paper chromatographic method. However, each of these substances was further purified by paper chromatography after a sufficient number of identical fractions from several chromatograms had been accumulated to ensure that a significant improvement in purity could be achieved. Ultimately, each fraction was obtained in purity exceeding 90% and the infrared spectrum was measured (KBr pellet). The spectra of the five components are given in Figs. 1–5 from which it is evident that components 1, 4, and 5 differ markedly from clam poison. Component 2, however, has a spectrum identical with that of clam poison dihydrochloride except that a weak band present in the toxin at 1770 cm^{-1} is absent from component 2. It is concluded, therefore, that the two substances are very similar in structure. The infrared spectrum of clam poison dihydrochloride is not very helpful from the viewpoint of identification of structural features of the molecule. Schantz and co-workers (15) have pointed out that the spectrum suggests the presence of OH or NH groups (3100–3300 cm^{-1}); C=O or C=N linkages (1600–1750 cm^{-1}) with a good possibility of the presence of an amide structure; and C—O or C—N bonds (1100 cm^{-1}). Also, the absence of a lactone, ester, anhydride, or aromatic structure seems assured. The only addition which we can make to this evaluation is the suggestion that the bands in the 1680–1770 cm^{-1} region could owe their origin to the presence of a cyclic imide, cyclic urea, or hydantoin-type structure (16, 17, 18).

Components 1, 2, and 5 when chromatographed again under the same conditions by which they were first isolated, each gave only one spot of R_F identical with that originally found. These substances were concluded to be impurities in the original sample, since a purer reference specimen of clam poison dihydrochloride furnished later by the U.S. Army Chemical Warfare Laboratories did not contain any of these materials.

Chromatography of pure clam poison dihydrochloride (component 3) on Whatman No. 1 paper which had been pretreated according to procedure 1 or 2 (Table II) gave only one Weber-positive spot, R_F 0.32, but the use of *unwashed* Whatman No. 1 paper under otherwise identical conditions led to two Weber-positive spots, R_F 0.32 and 0.53. Repeated recovery and rechromatography of the substance of R_F 0.32 on *unwashed* paper consistently gave the same result. Two spots of identical R_F were also obtained when clam poison was chromatographed on pretreated or *unwashed* paper after basification with ammonia. Component 4 when rechromatographed on *unwashed* or pretreated Whatman No. 1 paper gave two Weber-positive spots, R_F 0.32 and 0.53, with the latter predominating unless the material was acidified with hydrochloric acid prior to chromatography, when only one Weber-positive spot, R_F 0.32, was obtained. Multiple spots have



FIGS. 1–5. Infrared spectra of clam poison components 1–5.

frequently been observed during chromatography of pure substances and a review of the causes has been given by Pollard *et al.* (19). Where polyfunctional basic compounds are chromatographed in acidic solvent systems, the explanation most frequently invoked is the generation of more than one ionic species of the base by reaction of the substance with the mobile phase (20, 21), but the formation of tautomers has also been postulated (22). Since clam poison has two basic centers and Mold *et al.* (1) obtained some evidence from countercurrent distribution studies at pH 8 that clam poison consists of two tautomers of differing toxicity, either of the explanations mentioned above was a possibility.

Owing to the marked similarity in chromatographic behavior of the various materials of R_F 0.32 and 0.53 described in the foregoing paragraph it was initially assumed that all the substances of identical R_F were identical and that the substance of R_F 0.32 was an ionic species of the toxin corresponding to a diacid salt whereas the substance of R_F 0.53 was an ionic species corresponding to a monoacid salt. Such behavior is exhibited by certain alkaloids of the lupin group where stable monoacid and diacid salts have been obtained (23, 24). Some support for this view was obtained by the behavior of clam poison acetate on paper chromatography which is now to be described. The acetate was prepared from the dihydrochloride by ion exchange on Amberlite IRA-400 resin and gave an infrared spectrum (Fig. 6) which differed from that of the dihydrochloride only in those

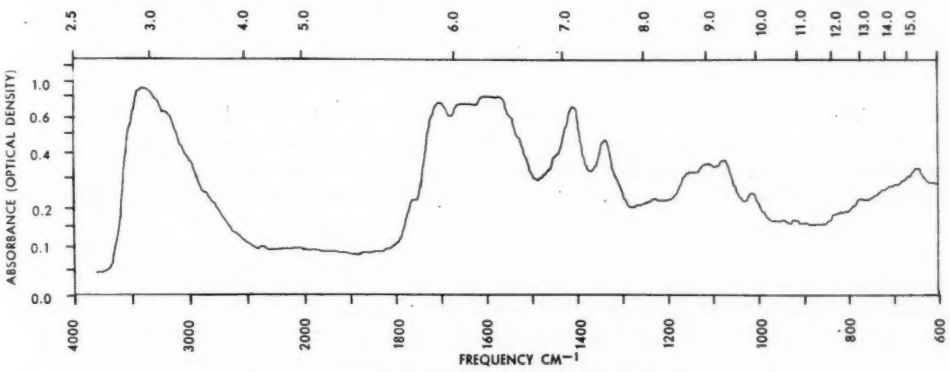


FIG. 6. Infrared spectrum of clam poison acetate.

regions where the ionized carboxyl group would be expected to modify it. Paper chromatography of the acetate on *unwashed* Whatman No. 1 paper gave only one Weber-positive spot, R_F 0.53, but on paper pretreated according to procedure 1 or 2 (Table II) a faint spot at R_F 0.32 was observed in addition to the strong one at 0.53. Reconversion of the acetate to the hydrochloride via ion exchange on Amberlite IRA-400 resin gave a substance with infrared spectrum identical with that of the original dihydrochloride and which showed no abnormality in paper chromatographic behavior (i.e. one spot on pretreated paper, R_F 0.32, two spots on *unwashed* paper, R_F 0.32 and 0.53). These results, coupled with the observation that the major component of clam poison and chloride ion have identical R_F's on *unwashed* paper suggested that the components of R_F 0.32 were the dihydrochloride, whereas the components of R_F 0.53 were the acetate possibly in the mono form. Unfortunately, however, this unifying theory had to be rejected in the light of the work now to be described.

If the substances of R_F 0.32 and 0.53 were different ionic species of the same base in

equilibrium, then it should be possible to convert one to the other by a suitable adjustment of pH, and the same situation could obtain in the event that the two substances were tautomers in equilibrium. Judging from the chromatographic behavior described above, high pH's favor the faster-moving species and low pH's favor the slower-moving species. Even if there were a marked difference in toxicity between the two, it should be possible to regenerate the toxicity of pure clam poison by acidification to pH 3, which is the pH of maximum stability of clam poison and therefore the pH at which the bio-assay is performed (13). The complete lack of toxicity of component 4 (see Table IV) gave an indication that the simple ionic species theory is inadequate and furnished definite evidence that component 4 is not the minor tautomeric constituent B isolated by Mold *et al.* (1). The latter substance possessed significant toxicity but less than that exhibited by the major constituent A. It was observed furthermore (1) that if constituent B was acidified and allowed to stand for 16–24 hours the toxicity finally observed was identical with that of constituent A. Freshly prepared solutions of component 4 in hydrochloric acid of pH 3.0 showed no toxicity at a dosage of 1 mg/kg and no toxicity developed if such solutions were kept for periods of up to 5 days at pH 3.0 or 1.3. Chromatography of these solutions on pretreated Whatman No. 1 paper gave one Weber-positive substance, R_F 0.30, quantitative recovery and bio-assay of which confirmed that it was not clam poison dihydrochloride, since it was non-toxic. It was therefore concluded that component 4 is an impurity in, or a decomposition product from, the original clam poison sample, rather than another ionic species of the toxin or the tautomeric constituent B isolated by Mold *et al.* (1). Furthermore, the Weber-positive substance of R_F 0.30 generated by the action of aqueous hydrochloric acid on component 4 is not clam poison.

Attention was then turned to the substance of R_F 0.53 obtained by chromatography of pure clam poison on *unwashed* Whatman No. 1 paper, which is hereinafter referred to as substance X. An attempt was made to establish that substance X was identical with clam poison acetate. The substance X obtained by chromatography of the pure dihydrochloride on *unwashed* paper was impure because of the relatively large quantity of extractable impurity present in Whatman No. 1 paper. Rechromatography on specially pretreated (procedure 2) paper effected only a partial purification so that positive identification via the infrared spectrum was prevented. The substance possessed sufficient toxicity, however (525 MU/mg), to rule out the possibility that it is identical with component 4 from the original toxin sample. The substance was converted to the chloride salt by ion exchange in the usual manner, in the expectation that, after further purification by chromatography, a specimen of clam poison dihydrochloride would be obtained which could be identified by its infrared spectrum. A Weber-positive substance, R_F 0.35, was obtained which had a toxicity of 2150 MU/mg, but the infrared spectrum (Fig. 7) was quite different from that of clam poison dihydrochloride. It is therefore concluded that component 4, substance X, and clam poison acetate are non-identical and that the first two of these are not merely different ionic species of clam poison. It is also no longer considered that the substance of R_F 0.53 generated by the action of ammonia on clam poison is identical with component 4, substance X, or clam poison acetate. The nature of substance X and its relation to clam poison is still obscure but it seems likely that it is most probably a degradation product. If such is the case, it is of considerable interest because of the mild conditions of its formation and its intrinsic toxicity. A more detailed investigation of this substance must await more effective means for its purification. Clam poison acetate was shown to be a diacid rather than a monoacid salt by a pK determination, which revealed the presence of two basic groups pK_1 8.1 and pK_2 11.1 and a molecular weight in agreement with that reported by Schantz *et al.* (2) for the dihydrochloride.

To summarize the results of the multiple spots observed during chromatography of the samples of clam poison dihydrochloride and acetate available to us it has been concluded that component 4 from the original sample is a non-toxic impurity which is not reconverted to clam poison by treatment with acid. It is not identical with clam poison acetate or substance X, which is formed as a minor Weber-positive component from clam poison dihydrochloride during chromatography on *unwashed* Whatman No. 1 paper. Substance X is a toxic unidentified substance which is possibly a degradation product of clam poison. It also does not regenerate clam poison on treatment with acid but does produce another toxic Weber-positive substance of R_F 0.35. Clam poison acetate is a diacid salt like the hydrochloride and is convertible to the latter by ion exchange. The substance of R_F 0.53 formed during chromatography of clam poison dihydrochloride after basification with ammonia is not considered to be identical with component 4, substance X, or clam poison acetate.

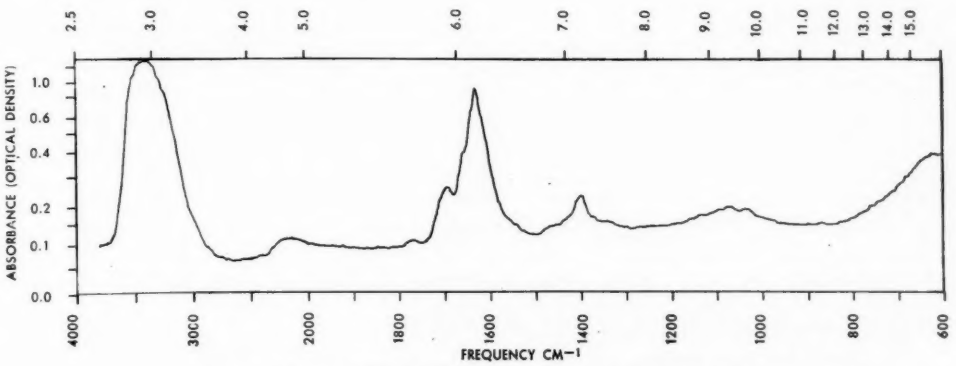


FIG. 7. Infrared spectrum of substance X in the chloride form.

The usefulness of the paper chromatographic method for purification of clam poison tailings has been studied. Preliminary chromatograms with the acetate revealed the presence of three ultraviolet-fluorescent spots, R_F 0.51, 0.53, and 0.62, the last two of which were Weber-positive. As expected, the major component had R_F 0.53. The chloride form, on the other hand, gave on chromatography five ultraviolet-fluorescent spots, R_F 0.20, 0.25, 0.30, 0.36, and 0.58, the last three of which were Weber-positive. The major component had R_F 0.30. It was therefore evident that better resolution of the Weber-positive components and therefore a better purification of the tailings was to be expected by the use of the chloride salt.

A sample of clam poison tailings (11.7 mg, 2000 MU/mg) was purified by quantitative chromatography on an $18\frac{1}{4}\times22\frac{1}{4}$-in. sheet of specially pretreated (procedure 2) Whatman No. 1 paper. The results given in Table V show that the percentage recovery on a weight basis was 95.5% and on a mouse unit assay basis was 81.5%. As expected, most of the toxicity (76%) was contained in the Weber-positive fraction of R_F 0.30, and a substantial improvement in the purity of the toxin was achieved judging by the enhanced toxicity. A portion of the recovered toxin (5.30 mg) was repurified by the same method, yielding 3.26 mg of clam poison dihydrochloride with toxicity 4150 MU/mg. Thus, 13,500 MU of a possible 14,500 MU (93%) was recovered in the second treatment. The total recovery for two purifications was 70.5% and the purity was increased from ca. 36% to ca. 75% (assuming the toxicity of the pure hydrochloride to be 5500 MU/mg). The infrared

spectrum of the final product was identical with that of clam poison dihydrochloride. These results clearly establish the usefulness of the paper chromatographic technique as a method for enrichment of clam poison tailings on a small scale. In contrast to this, paper or cellulose column chromatography using 95% ethanol, triethylamine (1:1) as developer, chromatography on Amberlite XE-64 resin, and chromatography in absolute ethanol on Woelm's acidic alumina was found to be incapable of improving the purity of clam poison tailings (4). Paper chromatographic purification of the tailings in the acetate form was also briefly examined. The toxic component was located, as expected, at R_F 0.53. The enrichment was not quite as great as in the case of the hydrochloride, the toxicity on one purification being increased to 2200 MU/mg from 1600 MU/mg.

TABLE V

Purification of clam poison dihydrochloride tailings

Fraction No.	R_F	Corrected weight (mg)	Toxicity (MU/mg)	Total MU
UV_1	0.20	0.37	400	148
UV_2	0.25	1.81	400	725
W_1	0.30	5.78	2900	16800
W_2	0.36	0.98	167	163
W_3	0.58	1.62	93	151
		10.56*		17987*

*Theoretical recovery = 11.05 mg or 22,100 MU.

EXPERIMENTAL

Paper Chromatography of Clam Poison Dihydrochloride

For paper chromatography, 20 microliters of a 1.2% aqueous solution (pH 5) (240 γ) of the clam poison dihydrochloride sample received from Schantz *et al.* (2) was applied to *unwashed* Whatman No. 1 paper and developed overnight in a glass chromatographic tank by the descending technique. The air-dried chromatograms were sprayed with Weber reagent (3), Sakaguchi reagent (5), or bromcresol green indicator, or examined under an ultraviolet lamp for location of the components. Results are given in Table I for several solvent systems.

Procedure for Pretreatment of Papers

All papers used were Whatman No. 1, $18\frac{1}{4}\times22\frac{1}{4}$ in. on which $\frac{1}{2}$-in. serrations were cut along the bottom edge to assure even dripping during washing. The papers were washed chromatographically in a Model A 300 Chromatocab (Research Equipment Corp., Oakland, California) by the descending technique. Between each washing step, the papers were air-dried at room temperature. The washing procedures examined are shown in Table II.

Blank Runs on Pretreated Papers

The pretreated paper was developed overnight by the descending method until the solvent front was ca. 1 to $1\frac{1}{2}$ inch from the serrated edge (16 hours). The "chromatogram" was air-dried at room temperature and a vertical strip $1\frac{3}{8}$ in. wide (corresponding to the reference strip removed for exact location of the Weber-positive spots in actual chromatograms) was removed. Rectangular strips were then cut from the remainder of the chromatogram at R_F's corresponding to those of the clam poison components. Since the

widths of the bands formed in actual chromatograms were found to vary, the widths of strips removed for extraction in the control runs were not identical for all components and were as follows: components 1, 2, and 4, 1 in.; component 3, $1\frac{1}{8}$ in.; component 5, $\frac{3}{4}$ in. The strips were eluted with water (1.5–2.0 ml) and the eluates transferred to small tared flasks with ground caps which had been previously weighed on a microbalance. The solutions were frozen, lyophilized, and the residues were weighed again. The quantities of residues obtained corresponding to the various components of the clam poison sample from papers pretreated according to the methods given in Table II are shown in Table III.

Purification of Clam Poison Dihydrochloride

A sample of the partially purified clam poison dihydrochloride solution (1.2% aqueous) containing 13–17 mg total solids was applied to a sheet of Whatman No. 1 paper ($18\frac{1}{4} \times 22\frac{1}{4}$ in. pretreated according to procedure 2) as 18 spots located on a start line 3 inches from the top of the paper. The two outside spots were placed $\frac{3}{4}$ inch from the edge of the paper and all others were spaced approximately 15/16 inch apart. To keep the area of the applied spots small, 5λ aliquots of the solution were applied at each location, and all spots were allowed to dry prior to subsequent application. The paper was developed exactly as described for the control runs. The air-dried chromatogram was examined under ultraviolet light to determine the straightness of the bands, and the width of component 1. A vertical strip $1\frac{1}{8}$ in. wide was cut from the chromatogram and sprayed with Weber and Sakaguchi reagent to determine the exact widths of the bands due to components 2, 3, 4, and 5. The bands were cut from the chromatograms and eluted with water (1.5–2.0 ml). Each eluate was passed through IRA-400 (Cl) resin (1 ml) to reconvert the basic components to the chloride form. A fresh sample of resin was used for each component and each column was eluted with a further 2 ml of water to ensure complete recovery. The resin used in these experiments was washed prior to use for 1 week with water after conversion to the chloride form, and it was ascertained in independent experiments that similar treatment of eluates from blank runs introduced no significant amounts of impurities. The solutions were transferred to tared flasks, and the components recovered and weighed as described for the blank runs. The areas of the strips of paper from which the components were eluted were measured by means of a planimeter, and the weight of impurity introduced during elution of each strip calculated after the data in Table III was converted to milligrams of impurity per square inch of paper extracted. The weights of impurities thus calculated were deducted from the weights of components recovered by the procedure given above, to give the actual weight recovered. The weight of material lost by spraying the reference strip was deducted from the original sample weight to give the actual weight of sample used. Results for a typical assay by this method are shown in Table IV. Clam poison dihydrochloride (component 3) was obtained as a colorless to pale yellow mixture of glass and crystals which did not melt but sintered above 220°, $[\alpha]_D^{25}$ +128° (1% aqueous solution). The other components of the sample were obtained as colorless to pale yellow glasses. Bio-assays were performed by the mouse unit (MU) assay method of Schantz *et al.* (13) and infrared spectra were measured as potassium bromide pellets (0.5 mg of sample + 300 mg of KBr) on a Perkin–Elmer model 21 spectrometer.

Attempted Reconversion of Component 4 to Clam Poison Dihydrochloride

Component 4 recovered from the purification of clam poison dihydrochloride was dissolved in hydrochloric acid at pH 3.00 at 20°, producing a solution containing exactly 200 γ/ml. Aliquots of this solution were removed and solutions containing 1, 5, 10, and

20 γ/ml were prepared using the dilute hydrochloric acid mentioned above. Bio-assays of these solutions by the MU assay method revealed that none of these solutions was toxic at dosage levels of 50–1000 γ/kg when freshly prepared, nor did toxicity develop in such solutions on standing for periods of up to 5 days at room temperature. The experiment was repeated using hydrochloric acid of pH 1.30 for preparation of the solutions but the increased acidity did not change the results in any respect. Chromatograms performed on the residual solutions from the experiments described above, using pretreated (procedure 1) Whatman No. 1 paper gave one Weber-positive spot, R_F 0.30. Quantitative isolation of the latter material from a separate chromatogram followed by bio-assay confirmed its non-toxic nature.

Isolation of Substance X *and Attempted Reconversion to Clam Poison Dihydrochloride*

A sample of clam poison dihydrochloride (24.2 mg) was chromatographed quantitatively using the procedure described previously for the purification of the toxin, except that two sheets of *unwashed* Whatman No. 1 paper were used. Two control chromatograms were run simultaneously to determine the quantities of extractables recovered at R_F's corresponding to the two Weber-positive substances formed. In this manner 21.2 mg of clam poison dihydrochloride (R_F 0.32) containing 2.8 mg of impurity and 15.1 mg of substance X (R_F 0.53) containing 11.8 mg of impurity were obtained. Substance X was thus only ca. 20% pure and was rechromatographed quantitatively on pretreated (procedure 2) Whatman No. 1 paper. After two such treatments, the purity of substance X had been increased to only 33%. An infrared spectrum on a portion of this material therefore gave an inconclusive result (i.e. identity or non-identity of substance X with the Weber-positive material of R_F 0.53 obtained during the paper chromatography of clam poison acetate was not established) but bio-assay revealed that substance X, even in its impure condition, had a toxicity of 525 MU/mg. The substance X remaining (6.50 mg, containing 2.24 mg of pure substance X) was passed through an Amberlite IRA-400 (Cl) resin column and the resultant material was purified by quantitative paper chromatography on pretreated (procedure 2) Whatman No. 1 paper, yielding 1.32 mg of Weber-positive substance, R_F 0.35, of estimated purity 75%. This substance had an infrared spectrum (Fig. 7) quite different from that of clam poison dihydrochloride and had a toxicity of 2150 MU/mg. Clam poison dihydrochloride of the same purity (recovered from tailings) had a toxicity of 4150 MU/mg and an infrared spectrum identical with that of a pure specimen of the dihydrochloride.

Purification of Clam Poison Tailings

A solution of clam poison tailings in the acetate form (220 ml, 1.16 mg/ml total solids, toxicity 1600 MU/mg) from the column chromatographic purification method of Schantz *et al.* (2) was found to be too dilute for paper chromatographic work and was lyophilized, yielding 255 mg of residue. The residue was dissolved in water to give a solution containing 25.5 mg/ml of total solids. Exactly 5 ml of the solution was converted to the chloride form by elution through Amberlite IRA-400 (Cl) resin, the eluate was lyophilized, and the residue (108 mg) made up to exactly 5 ml in water. A bio-assay on the two solutions revealed the acetate to have toxicity 1600 MU/mg and the chloride 2000 MU/mg.

Clam poison tailings chloride (11.7 mg) was purified on an $18\frac{1}{4}\times22\frac{1}{4}$-in. sheet of pretreated (procedure 2) Whatman No. 1 paper in the same manner as described for the original clam poison sample, after a control chromatogram had been performed to determine the weight of impurity which would be introduced into each fraction. A bio-assay

was performed on each fraction and results are given in Table V. Fraction W_1, R_F 0.30, accounted for 76% of the toxicity of the original sample and exhibited enhanced toxicity (2900 MU/mg). A portion of fraction W_1 (5.30 mg) was repurified by the same method, using, however, a 12×18¼-in. sheet of pretreated paper, and gave 3.26 mg (93%) of clam poison with toxicity 4150 MU/mg. The infrared spectrum of the final product was identical with that of clam poison dihydrochloride.

Chromatographic Behavior of Clam Poison Acetate

Clam poison acetate was prepared from the dihydrochloride by ion exchange using Amberlite IRA-400 resin. The infrared spectrum of the acetate was measured and is shown in Fig. 6. Reconversion of a sample of the acetate to the dihydrochloride by ion exchange produced material with an infrared spectrum identical with that of the original dihydrochloride.

Paper chromatograms with the acetate on *unwashed* Whatman No. 1 paper in the usual manner gave only one Weber-positive spot, R_F 0.53, but on paper pretreated according to procedures 1 or 2 (Table II) a faint spot at R_F 0.30 was observed in addition to the strong one at 0.53.

Potentiometric Titration of Clam Poison Acetate

For potentiometric titration ca. 5-mg samples of clam poison acetate were dissolved in CO_2-free water in a nitrogen-swept microflask equipped with a Beckmann Probe Assembly 290-32 glass electrode and 270-32 reference electrode (attached to a Beckmann Model GS pH meter) and Gilmont ultramicroburet, and titrated vs. 0.1 *N* barium hydroxide. Found: pK_1 8.1, pK_2 11.1, molecular weight 455. Calc. for $C_{10}H_{17}N_7O_4 . 2\ CH_3COOH$, molecular weight 419.

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COMPLEX MANGANESE PERIODATES[1]

M. W. Lister and Y. Yoshino

ABSTRACT

The preparation of complex periodatomanganates, from manganous salts reacting with sodium (or potassium) hypochlorite and periodate, is described. The products were $Na_7H_4Mn(IO_6)_3.17H_2O$ and $K_7H_4Mn(IO_6)_3.8H_2O$. The manganese was shown to be tetravalent from its redox equivalent in its reaction with iodide and with ferrous ions, and from its magnetic moment, which was 3.87 Bohr magnetons, corresponding to three unpaired electron spins. The ion has an absorption maximum at 483 mμ. Rather complicated pH curves were found for this compound in acid–base titrations. The complex ion decomposes slowly to permanganate and iodate; the kinetics of this reaction have been measured, and a possible mechanism has been suggested. The reaction proceeds in several steps by way of simpler complex ions. The same applies to the reaction between iodide ions and the complex ions when the pH is about 9.

In a recent paper (1) the present writers reported on the preparation and properties of complex cobalt periodates. Similar copper compounds have been prepared and examined (2, 3), as have somewhat similar nickel compounds (4). The ability of periodate ions to oxidize manganese to permanganate suggested that manganese would not give a similar compound; but, on making the experiment, it was found that complex salts of tetravalent manganese could in fact be prepared. The object of the present paper is to describe the preparation and properties of these compounds.

1. *Preparation of Compounds*

(a) *Sodium Periodatomanganate* (IV)

After various preliminary attempts, the method used in a typical preparation was as follows. Twenty grams of sodium periodate, $Na_2H_3IO_6$, was dissolved in a mixture of 200 ml of water and 25 ml of 6 *M* nitric acid. To this was added 5 g of manganous chloride tetrahydrate dissolved in 20 ml of water, and then, before any precipitation could occur, a mixture of 100 ml of 1.5 *M* sodium hypochlorite and 15 g of sodium hydroxide dissolved in 20 ml of water. Some brown precipitate formed (presumably manganese dioxide), and after centrifuging this off, a red solution was obtained. Alcohol was added till the solution was 40% alcohol by volume, when a red sticky material precipitated. This was dissolved in 100 ml of dilute sodium hydroxide, filtered on sintered glass, and reprecipitated with alcohol, when a dark red precipitate was obtained. This purification process was repeated twice, and the precipitate was finally dissolved in water, which was allowed to evaporate slowly. Red crystals were obtained, filtered off, washed with a little cold water, and dried. The yield was somewhat variable, but was between 4.5 and 10 g on various occasions.

To analyze the solid, it was dissolved in water and sulphur dioxide was passed into the solution; complete reduction occurred. Excess sulphur dioxide was boiled out, and silver nitrate was added. Silver iodide was filtered off, washed, dried, and weighed. The silver in the filtrate was removed as silver chloride, and the solution was run through a column of Dowex 50W-X8 cation exchange resin in its hydrogen form. The sodium and manganese ions were adsorbed, and the sodium was then eluted with 0.4 *M* HCl, which leaves the manganese on the column. The sodium solution was evaporated and the sodium chloride was weighed; the purity of the sodium chloride was then checked

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by silver nitrate titration. The column was then further eluted with 2.0 *M* HCl, which removes manganese. Sulphuric acid was added to the manganese solution, which was evaporated, and the manganese was weighed as $MnSO_4$. The water content of the crystals was determined from the weight loss at 150° C.

The results were Na 13.3%, Mn 4.60%, I 31.6%, water 25.3%. Calculated for $Na_7H_4Mn(IO_6)_3.17H_2O$ are Na 13.4$_5$%, Mn 4.60%, I 31.8$_5$%, water 25.6%. This analysis is the mean of two closely agreeing analyses.

A confirmation of this formula was obtained by passing a solution of a known weight of sodium periodatomanganate (IV) through an anion exchange resin (Dowex 1-X4) in its chloride form. The displaced chloride was titrated with silver nitrate. It was found that 1.415×10^{-3} g-molecules of chloride was displaced by 0.2016×10^{-3} g-molecules of the sodium salt, a ratio of 7 to 1. If R is the resin, the reaction is

$$7RCl + Na_7H_4Mn(IO_6)_3 \rightarrow R_7H_4Mn(IO_6)_3 + 7NaCl.$$

(b) Potassium Periodatomanganate (IV)

This compound was prepared in a similar fashion to the sodium salt, but with potassium periodate, KIO_4, and potassium hypochlorite as starting materials. It proved to be much less easy to crystallize than the sodium salt, as it is evidently more soluble in water, and readily gives a rather sticky precipitate with alcohol. An alternative preparation by running a solution of the sodium salt through a cation exchange resin, in its potassium form, gave a similar product. Slow evaporation over calcium chloride gave a dark red-brown solid.

Similar analysis as for the sodium salt gave K 23.3%, Mn 4.72%, I 32.4%, and H_2O 11.9%. Calculated for $K_7H_4Mn(IO_6)_3.8H_2O$ are K 23.9%, Mn 4.79%, I 33.2%, H_2O 12.6%.

(c) Barium Periodatomanganate (IV)

A solution of the sodium complex salt was mixed with barium hydroxide, when a brown precipitate formed. Its composition was approximately $Ba_5HMn(IO_6)_3.10H_2O$, but was somewhat variable, and no method was found to recrystallize the product, as it was insoluble in water and decomposed by acids. Its observed oxidation equivalent (see below) with potassium iodide and acid was 63.2 (calculated 61.2). It is probably the above salt contaminated by impurities such as barium carbonate, or barium periodate.

2. Valency of Manganese

The analyses above do not, of course, give any indication of the valency of the manganese. This was determined by various titrations, and by magnetic measurements, which are now described.

(a) Iodometric Titration

A weighed amount of the sodium salt was added to acidified potassium iodide. The liberated iodine was titrated with thiosulphate. The observed equivalent weight varied from 45.66 to 45.75 in three analyses, with a mean of 45.72. The calculated equivalents are Mn (III) 47.84, Mn (IV) 45.95, Mn (V) 44.22; these assume that the periodate is reduced to iodine and the manganese to Mn^{++}. The results agree with the presence of manganese (IV) in the compound.

(b) Potentiometric Titration with Ferrous Ammonium Sulphate

It was found that sodium periodatomanganate, like the corresponding cobalt salt, could be titrated against ferrous ions in acid solution. A weighed amount of the salt

was dissolved in 1 *N* sulphuric acid and titrated against approximately 0.1 *M* ferrous ammonium sulphate. The titration was followed potentiometrically with the cell: $Pt/Na_7H_4Mn(IO_6)_3/Hg_2Cl_2$ in saturated KCl/Hg. The platinum was the positive pole, and a typical titration of 0.0037 *M* complex salt against 0.0995 *M* Fe^{++} at 25° C gave:

ratio										
0	1.97	3.94	5.90	7.38	7.87	7.92	7.97	8.06	8.36	8.86
e.m.f. (volts)										
1.102	1.098	1.090	1.089	1.082	1.045	0.750	0.711	0.653	0.578	0.542

Here "ratio" means the number of gram-molecules of ferrous ions divided by the gram-molecules of complex salt. In this titration the inflection comes at 7.90; in another titration this was at 7.99; it is evident that 1 g-molecule of the complex salt reacts with 8 of ferrous sulphate.

$$2Na_7H_4Mn(IO_6)_3 + 16FeSO_4 + 17H_2SO_4 \rightarrow 7Na_2SO_4 + 2MnSO_4 + 8Fe_2(SO_4)_3 + 6HIO_3 + 18H_2O.$$

This also supports a formula with tetravalent manganese. With the corresponding cobalt compound, and with sodium periodate itself, there is reason to believe that the cell is not giving the true equilibrium voltage (see reference 1); this probably applies to the sodium periodatomanganate also. However, the evidence that this is tetravalent manganese is quite definite.

(c) Reduction with Sulphur Dioxide

An aqueous solution of the sodium periodatomanganate is reduced by sulphur dioxide, and after the excess sulphur dioxide was boiled off, the sulphuric acid produced in the oxidation can be titrated with sodium hydroxide. One gram equivalent of sulphuric acid was produced by 60.1 g of the compound; calculated are Mn (III) 62.9, Mn (IV) 59.7, Mn (V) 56.8. Thus this result also supports a formula with Mn (IV). The reaction is

$$Na_7H_4Mn(IO_6)_2 + 13SO_2 + 8H_2O \rightarrow 3NaI + 2Na_2SO_4 + MnSO_4 + 10H_2SO_4.$$

(d) Magnetic Properties

Magnetic measurements were made on the solid compound, using a Goüy apparatus of the conventional type. The compound was suspended in a tube surrounded by a jacket in which water at various temperatures could be circulated. The tube was calibrated with sucrose, whose susceptibility was taken to be -0.566×10^{-6}/g (5). The observed susceptibilities are given below; as is to be expected the compound is paramagnetic, and the magnetic moment was calculated after allowing for the diamagnetic susceptibility of the various atoms by means of Angus's method (6).

Temp., °C	3.0	25.6	39.7
χ_g	4.695	4.323	4.154×10^{-6}
$\chi_{g\text{-molecule}}$	5611	5166	4964×10^{-6}
Corr. for diamagnetism	6141	5696	5494×10^{-6}

These results fit a Curie–Weiss equation, $\chi = C/(T+\theta)$, where $\theta = 21°$ and $C = 1.826$. This makes the magnetic moment 3.84 Bohr magnetons. The calculated value for a "spin-only" atom with three unpaired electron spins is 3.87. Hence these results strongly support a formula with Mn (IV).

3. Absorption Spectrum

Solutions of sodium periodatomanganate (IV) are red in color. The absorption spectrum

measured on a Beckmann DU spectrophotometer was as follows for a 1.04×10^{-3} M solution containing 0.100 M sodium hydroxide:

wavelength (mμ)									
800	750	700	650	600	580	560	540	520	500
490	480	470	460	450	440	430	420	410	400
extinction coefficient									
3	4	7	13	54	98	169	266	354	409
423	428	417	398	378	369	393	453	602	870

There is thus a single maximum at 483 mμ. In acid solution the complex is stable enough to be measured; the general shape of the curve is the same but the maximum is less well marked. In 0.106 M perchloric acid, the spectrum in the neighborhood of the maximum (at 479 mμ) was:

wavelength	500	490	480	470	460	450
extinction coefficient	397	406	416	408	418	438

If the free acid was made by passing a solution of the sodium salt through an ion exchange resin in its hydrogen form, a somewhat different spectrum was obtained, with a shoulder at 480 mμ instead of a maximum. When the solution was allowed to stand for 2–3 days the spectrum of the permanganate ion appeared, but this was not visible at first. If the free acid was made alkaline with sodium hydroxide (0.05 M), a modified spectrum was obtained, again with a shoulder instead of a maximum at 480 mμ, though the general shape and size of extinction coefficients (412 at 480 mμ) were much the same. The nature of the change in the complex is not entirely clear, but related observations will be given below. If the adsorption maximum at 483 mμ is due to a $t_g \rightarrow e_g$ transition for one 3d electron, as seems possible, this corresponds to a splitting of 59 kcal between these levels. Sodium periodatocobaltate (III), which is diamagnetic, has a similar spectrum (1) with a maximum at 625 mμ.

4. pH Measurements

Since sodium periodatomanganate (IV) was fairly stable in acid solutions, its behavior in acid–base titrations was examined. The pH in these titrations was followed by a Beckmann pH meter with glass and saturated calomel electrodes. When 2.53×10^{-3} M sodium periodatomanganate (IV) was titrated against 0.1009 M perchloric acid, only one inflection was observed in the curve of pH plotted against the ratio of acid to sodium periodatomanganate (IV). This was after addition of two equivalents of acid, and corresponds to $Na_5H_6Mn(IO_6)_3$; the pH was then 7.48. These results are plotted in Fig. 1. If the acidified solution was allowed to stand briefly, and then titrated in the

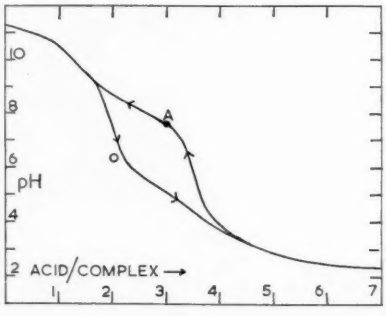


FIG. 1.

reverse direction with sodium hydroxide, a different curve resulted with an inflection at 4NaOH to one molecule of $H_{11}Mn(IO_6)_3$, that is, at $Na_4H_7Mn(IO_6)_3$. Thus the shape of the pH curves is like a hysteresis loop. Very nearly the same curve was obtained if the sodium periodatomanganate solution was run through a cation exchange resin in its acid form, and the resulting solution was titrated with NaOH; again the inflection was at $Na_4H_7Mn(IO_6)_3$. However, if the acid solution was allowed to stand for 24 hours before titration, the inflection shifted somewhat to about $3\frac{1}{2}$NaOH per molecule of the complex acid.

When 4NaOH was rapidly added per molecule of freshly prepared complex acid, the pH was 7.50 (A in Fig. 1); this is on the top branch of the loop. When a solution of $Na_7H_4Mn(IO_6)_3$ was prepared and $2HClO_4$ was rapidly added per molecule of complex, the pH was 6.32 on the lower branch of the loop; the same solution on standing for 24 hours had a pH of 7.68. If the free acid was made alkaline, its pH behavior was then the same as that of the original salt, so no permanent change was produced by acidification.

In another run the free complex acid was titrated against barium hydroxide. The pH was generally lower than in titrations with sodium hydroxide, but a precipitate formed. The first inflection on the pH curve was at $Ba_2H_7Mn(IO_6)_3$ (pH = 4.7), followed by a less well-defined rise between Ba_3 and Ba_4.

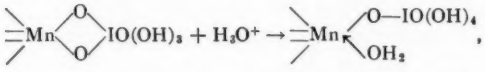
If the free acid was prepared by passing through a cation exchange resin in its acid form, and the free acid was immediately allowed to run through an anion exchange resin in its chloride form, three chlorides were displaced by one molecule of the complex acid, i.e.,

$$H_{11}Mn(IO_6)_3 + 3RCl \rightarrow R_3H_8Mn(IO_6)_3 + 3HCl.$$

However, the anion exchange resin acquired the brown color of finely divided manganese dioxide, and it is possible that all this experiment proves is that the sodium periodatomanganate (IV) contains three periodates per manganese. Each periodic acid displaced one chloride from the resin, as would be expected from its ionization constants, which are $pK_1 = 1.6$, $pK_2 = 7$ (7).

The pH of 2.5×10^{-3} M solutions of the free complex acid was found to be close to 2.30. This corresponds to two ionized hydrogens per molecule.

The explanation of all these observations is not too clear. Starting at the alkaline end, the main drop in pH was at $Na_5H_6Mn(IO_6)_3$ or $Na_5Mn(H_2IO_6)_3$. This is a formula with all periodates equally ionized. In the free acid two hydrogens are ionized, presumably as $(H_3O^+)_2[Mn(H_3IO_6)_3]^{--}$. The loop in the pH curve must be due to some rearrangement in the periodatomanganate ion on acidification, and this rearrangement, and its reverse on making alkaline again, must be slow. The nature of the process might be the replacement of a manganese–periodate bond by a band to water,



or dissociation to a lower complex (see later), but this is speculative. It is possible on this basis to write fairly plausible symmetrical formulae for the complex acid and its ions, but the evidence is not sufficient to establish them.

One other quantity can be calculated from these measurements. The pH of 0.00253 M sodium periodatomanganate was 10.87. Assuming that this is due to the hydrolysis of the $H_4Mn(IO_6)_3^{-7}$ ion, the hydrolysis constant can be calculated to be 3×10^{-4}.

5. Decomposition of Sodium Periodatomanganate (IV)

It was found that solutions of sodium periodatomanganate (IV) spontaneously changed on standing to a solution of sodium permanganate. The over-all reaction is

$$2Na_7H_4Mn(IO_6)_3 + 2H_2O \rightarrow 2NaMnO_4 + 3NaIO_3 + 3Na_2H_3IO_6 + 3NaOH.$$

Experiments were made on the rate of this reaction under various conditions. Solutions were kept immersed in a water thermostat, regulated to better than 0.1° C, and the change was followed by light adsorption measurements with a Beckmann DU spectrophotometer.

In the first experiments the pH was controlled by means of a sodium phosphate buffer, but when the rates were compared with those obtained in sodium hydroxide solution, it became evident that the buffer did not influence the rate solely through its control of pH. Presumably the phosphate altered the rate by combining with one of the reacting species. Hence the runs that will be considered below will be ones in which the reacting solution was made up solely from sodium hydroxide, sodium periodate, sodium periodatomanganate, and water. The optical density was mostly measured at 550 mμ, where the extinction coefficients are: $Na_7H_4Mn(IO_6)_3$ 217, Na_2MnO_4 768, $NaMnO_4$ 2190; this is an adsorption maximum for sodium permanganate. The results are given in Table I.

If the color density is assumed to be due solely to the complex and to permanganate, then we can calculate the complex concentration at different times. Runs 1, 2, and 3 show that the rate is proportional to the complex concentration, when other factors are kept the same. A plot of log (complex) against time is nearly linear for all three runs, with a slope of 2.80×10^{-3}. Runs 3, 8, 9, 10 give the effect of sodium hydroxide concentration. If $d\log(\text{complex})/dt$ is taken from a plot for each run the results are:

$$\begin{cases} d\log X/dt & 0.06 & 0.54 & 2.73 & 8.32\times10^{-4}\text{: X = [complex]} \\ [OH^-] & 0.05 & 0.02 & 0.01 & 0.005\,M \end{cases}$$

This is roughly, though not precisely, proportional to $[OH^-]^{-2}$; other powers of $[OH^-]$ give a much worse agreement with experiment. Finally runs 3, 4, 5, 6, and 7 give the effect of periodate concentration:

$$\begin{cases} d\log X/dt & 6.84 & 5.46 & 4.03 & 2.70 & 1.93\times10^{-3} \\ [\text{periodate}] & 2.46 & 3.27 & 4.91 & 7.37 & 9.82\times10^{-4}\,M \end{cases}$$

The rate is inversely proportional to the added periodate concentration, though there is some drift from this simple relationship at low concentrations. It is readily calculated that the extra hydroxide ions, contributed by the hydrolysis of the complex ion, should not alter the rate by more than ½% for any of these runs.

The temperature variation of the rate (runs 5, 11, 12) makes the apparent activation energy about 46 kcal. This is so high that it is evident it is not a genuine activation energy, but must contain other factors; these will be discussed below.

One other set of results deserves comment: these are runs 13, 14, and 15, all with no added periodate. Here again a plot of log (complex) against time is linear at first, though levelling off towards the end of the run. The slopes of the linear parts are not very different.

It is not possible to be sure of the detailed mechanism of the reaction from these rate measurements, although it must be somewhat as follows. The complex must presumably be in equilibrium with a simpler complex, which is the reactive species, and the inverse relationship to periodate and hydroxide concentrations shows the nature of this equilibrium. The products of decomposition of the reactive species are not so certain, but it

TABLE I

Decomposition of sodium periodatomanganate (IV)

Run 1 Temp. 35° C		Run 2 Temp. 35° C		Run 3 Temp. 35° C		Run 4 Temp. 35° C		Run 5 Temp. 35° C		Run 6 Temp. 35° C	
[Periodate] 7.37×10^{-4}	[OH^-] 0.01 *M*	[Periodate] 7.37×10^{-4}	[OH^-] 0.01	[Periodate] 7.37×10^{-4}	[OH^-] 0.01	[Periodate] 9.82×10^{-4}	[OH^-] 0.01	[Periodate] 4.91×10^{-4}	[OH^-] 0.01	[Periodate] 3.27×10^{-4}	[OH^-] 0.01
Time (min)	[Complex]	Time	[Complex]	Time	[Complex]	Time (min)	[Complex]	Time	[Complex]	Time	[Complex]
0	4.01×10^{-4} *M*	0	2.97×10^{-4}	0	2.16×10^{-4}	0	2.18×10^{-4}	0	2.22×10^{-4}	0	2.13×10^{-4}
7	3.91	9	2.87	10	2.09	10	2.15	9	2.11	5	2.07
14	3.78	17	2.77	20	1.97	20	2.08	19	1.95	10	1.95
24	3.54	26	2.61	35	1.80	30	1.98	30	1.74	18	1.77
34	3.28	42	2.35	55	1.55	55	1.79	45	1.50	27	1.57
45	3.05	56	2.14	85	1.27	82	1.59	67	1.20	37	1.38
60	2.77	75	1.90			116	1.35	94	0.96	55	1.10
92	2.27					150	1.16	100	0.92		
								134	0.71		

Run 7 Temp. 35° C		Run 8 Temp. 35° C		Run 9 Temp. 35° C		Run 10 Temp. 35° C		Run 11 Temp. 40° C		Run 12 Temp. 30° C	
[Periodate] 2.46×10^{-4}	[OH^-] 0.01	[Periodate] 7.37×10^{-4}	[OH^-] 0.05	[Periodate] 7.37×10^{-4}	[OH^-] 0.02	[Periodate] 7.37×10^{-4}	[OH^-] 0.005	[Periodate] 4.91×10^{-4}	[OH^-] 0.01	[Periodate] 4.91×10^{-4}	[OH^-] 0.01
Time (min)	[Complex]	Time	[Complex]	Time	[Complex]	Time (min)	[Complex]	Time	[Complex]	Time	[Complex]
0	2.05×10^{-4}	0	2.100×10^{-4}	0	2.04×10^{-4}	0	2.05×10^{-4}	0	2.06×10^{-4}	0	2.08×10^{-4}
5	1.90	11	2.095	10	2.03	6	1.87	8	1.64	12	2.05
12	1.73	27	2.084	32	1.99	12	1.67	16	1.27	38	1.96
20	1.50	74	2.079	93	1.86	23	1.34	24	0.99	63	1.84
26	1.36	135	2.059	132	1.76	30	1.16	35	0.73	103	1.64
35	1.16					33	1.08	48	0.55	145	1.44
50	0.93					43	0.90	72	0.34		
76	0.68					55	0.72	98	0.25		

Run 13 Temp. 35° C		Run 14 Temp. 35° C		Run 15 Temp. 35° C	
[Periodate] 0	[OH^-] 0.01	[Periodate] 0	[OH^-] 0.01	[Periodate] 0	[OH^-] 0.01
Time (min)	[Complex]	Time	[Complex]	Time	[Complex]
0	2.81×10^{-4}	0	1.97×10^{-4}	0	1.56×10^{-4}
4	2.55	5	1.68	5	1.33
9	2.18	10	1.44	6	1.30
15	1.85	15	1.23	10	1.13
23	1.51	20	1.06	16	0.98
35	1.16	25	0.93	26	0.79
		35	0.73	47	0.575
		55	0.49		

was observed that in partially reacted mixtures a weak adsorption band appeared in the neighborhood of $\lambda = 610$ to 620 mμ. This is probably due to manganate ions which have an adsorption maximum at 610 mμ. Manganate ions can be oxidized to permanganate by periodate, and some measurements have been made on this reaction which will be reported separately. In general two reaction schemes would fit the facts.

(A)

$$H_4Mn(IO_6)_3^{-7} + 2H_2O \rightleftharpoons H_3Mn(IO_6)_2^{-3} + H_3IO_6^{-2} + 2OH^- \quad [i]$$

$$H_3Mn(IO_6)_2^{-3} \rightarrow Mn(OH)_3IO_6^{-2} + IO_3^- \quad [ii]$$

$$Mn(OH)_3IO_6^{-2} + 2OH^- \rightarrow MnO_4^{--} + H_3IO_6^{--} + H_2O \quad [iii]$$

$$2MnO_4^{--} + H_3IO_6^{--} \rightarrow 2MnO_4^- + IO_3^- + 3OH^- \quad [iv]$$

Reaction [i] is a rapidly established equilibrium, and reaction [ii] is the slow stage. Reactions [iii] and [iv] would be at least fairly fast; reaction [iv] is not meant to be necessarily a simple single-stage reaction.

An alternative is:

(B)

$$H_4Mn(IO_6)_3^{-7} + 2H_2O \rightleftharpoons MnIO_6^- + 2H_3IO_6^{--} + 2OH^- \quad [i]$$

$$MnIO_6^- + H_3IO_6^{--} \rightarrow Mn(OH)_2IO_6^- + IO_3^- + H_2O \quad [ii]$$

$$Mn(OH)_2IO_6^- + 3OH^- \rightarrow MnO_4^{--} + H_3IO_6^{--} \quad [iii]$$

$$2MnO_4^{--} + H_3IO_6^{--} \rightarrow 2MnO_4^- + IO_3^- + 3OH^- \quad [iv]$$

As before, reaction [i] is a rapid equilibrium, and reaction [ii] is slow; reactions [iii] and [iv] are fairly fast. It will be recognized that for both schemes A and B the products of reaction [ii] might differ in detail, though the manganese must be hexavalent.

Applying ordinary kinetic considerations to scheme A, and assuming that the added periodate is not too low, we get the equation

$$\frac{2K.K_e}{3C_0[OH^-]^2} \cdot t = -\left(1+\frac{2u_0}{3C_0}\right) \ln \alpha - (1-\alpha)$$

where K_e is the equilibrium constant of reaction [i],
K is the rate constant of reaction [ii],
C_0 and u_0 are respectively the initial concentrations of complex and of added periodate,
and α is the fraction of complex left undecomposed.

If u_0 is not too small, then

$$\frac{d \ln (\text{complex})}{dt} = -\frac{2K.K_e}{[OH^-]^2\{2u_0+3C_0(1-\alpha)\}},$$

which will be reasonably constant at first. If an apparent rate constant, K^*, is defined by

$$\frac{d \ln (\text{complex})}{dt} = -\frac{K^*}{[OH^-]^2 u_0}$$

then $K^* = K.K_e/\{1+(3C_0/2u_0)(1-\alpha)\} = K.K_e$ initially. The observed values of K^* show a drift, with C_0/u_0, of the size predicted by the above equation, if the average value of α is taken as 0.8, which was in fact approximately correct. This makes $K^* = 2.1 \times 10^{-10}$ (g-molecules/l.)^3 min^{-1} at 35° C. The observed apparent activation

energy (46 kcal) will be the sum of the heat of reaction of reaction [i] and the activation energy of reaction [ii]; it is, of course, impossible to divide it between the two. It still seems to be rather high considering the nature of these two reactions.

The reactions of scheme B will give an identical equation relating α and t, if the same approximations are made. Hence the same remarks apply to scheme B.

Finally we must consider the observation that in the absence of added periodate a low, but fairly constant, value is obtained for d ln (complex)$/dt$. This could be explained for scheme A on the assumption that, under these conditions, the equilibrium of reaction [i] is fairly largely to the right. Then, at least initially, d ln (complex)$/dt = K\beta$, where β is the fraction dissociated. Runs 13, 14, and 15 make $K\beta$ about 3×10^{-2}. Allowing for the fact that $[OH^-] = 10^{-2}$ throughout, consistent results are obtained if we take $K_e = 2\times10^{-9}$ and $K = 0.11$. These numbers make the complex between 23 and 29% dissociated in these runs, and lead to a drift of the initial values d ln (complex)$/dt$ of about what is found. Thus although the results cannot give very accurate constants, they agree quite well with this explanation. Scheme B does not lead to a roughly constant value of d ln (complex)$/dt$ for different C_0 values, for the conditions of these runs, and hence scheme A is to be preferred.

It is also evident that the same dependence on hydroxide ion concentration cannot hold in acid solution, as, otherwise, acidified solutions would decompose much faster than they do. However, the pH titration curves gave evidence of rearrangement of the ion at moderate pH values, and this presumably also changes the rate of decomposition to permanganate.

One other set of experiments was made in the hope of throwing some light on the mechanism of decomposition. It was observed that when iodide was added to a solution of the complex, iodine was slowly liberated. It seemed possible that this might measure the rate of breaking up of the complex to a simpler ion and a periodate ion, which then rapidly oxidized iodide. If this were so, the rate should not depend on the iodide concentration. However, on making the experiment it was found that the rate at 25° C actually followed an equation

$$\text{rate} = -\frac{x}{x_0}\{0.9+3.0[I^-]\}10^{-6}$$

where x is the concentration of the complex at any time, and x_0 is its initial concentration; the time is measured in minutes. This is evidently a complex reaction, and the hope that it goes at a rate determined by the breaking up of the complex is not fulfilled. The inverse relation of rate to initial concentration of complex does presumably indicate an initial breaking up reaction, but it is not possible to suggest a detailed mechanism.

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SOME OBSERVATIONS ON THE THEORY OF π-BONDING IN INORGANIC COMPOUNDS[1]

A. D. WESTLAND

ABSTRACT

Evidence for the existence of multiple bonding due to *d*-electron donation by non-inert gas type ions is discussed briefly. A quantitative relation between the heats of formation of binary halides and chalcides containing non-inert gas type cations, on the one hand, and inert gas type cations, on the other, reveals an enhanced lattice energy in the former class of compounds. It is shown that this may be accounted for in the case of iron group elements without recourse to the postulate of participation of *d*-electrons in the bonding, as satisfactory estimates of the lattice energies of halides and sulphides may be made on the basis of values for group II*a* compounds.

The bond between the metal atom and carbon in carbonyl compounds is described as consisting of a dative σ-bond and a π-bond, the latter resulting from the overlap of an occupied *d*-orbital of the metal atom and a vacant orbital of the ligand (1). Such bonding is called upon to explain the observations of Carleson and Irving (2) and Ahrland and Chatt (3) that throughout a certain group of metals designated as "*b*-class" elements the stabilities of halo complexes increase in the order F < Cl < Br < I, whereas the order for most elements is I < Br < Cl $\ll$ F. Chatt and co-workers have extended the argument to complexes containing sulphur, phosphorus, etc., which elements they conceive as π-acceptor ligands (4), and they have proposed that the trans-directing influence of certain ligands in substitution reactions may be likewise ascribed to *d*-orbital participation (5).

The postulate of double bonding to the carbonyl group seems unavoidable if an implausibly large accumulation of negative charge at the metal atom is to be avoided. Apart from this rule of "small formal charge", evidence hitherto cited in support of π-bonding has been of an equivocal character. Thus Keller and Parry argue that certain cyano complexes such as $Cr(CN)_6^{3-}$ can hardly owe their great stability to π-bonding because they lack sufficient *d*-electron pairs (6). A recent argument (7) purporting to demonstrate π-bonding through a quantitative relation between absorption band frequencies and complex stability has been refuted by Williams, who maintained that these results and also the relative stabilities of halide complexes of *b*-class elements may be explained satisfactorily as a consequence of polarization of the ligands by the central atom (8). The *b*-class elements are just those which, as cations, have unusually large polarizing power due to inefficient shielding of the nuclear charge by their outer *d*-electrons. It may thus be argued that π-bonding is not required to explain many of the effects which have recently been attributed to it.

Williams' conclusions appear to be based upon a somewhat unsatisfactory argument, however. His measure of polarizing power, R = [ionization potential/$(100Z^2/r)$], is apparently meaningless for ions of inert gas configuration for it increases with increasing atomic number in a given periodic group. There is, however, little doubt that his relations correctly emphasize the fact that polarization may be used to explain many of the phenomena attributed to *d*-electron bonding. It seems that a thermochemical comparison of complexes containing central ions of inert gas and of non-inert gas configurations

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Contribution from the Department of Chemistry, University of Ottawa, Ottawa, Ontario.

respectively is desirable but this is not yet possible because of the lack of sufficient data for the former. Among binary compounds, however, it is possible to make such comparison and in what follows it is demonstrated that Williams' conclusions are borne out by the thermochemistry of transition metal halides and chalcides.

It has been shown in a previous publication (9) that it is possible to relate the heats of formation of various series of compounds with good precision by means of the quantity

$$-\Delta H^{**} = (-\Delta H_f + S + I)/\text{equivalents per mole},$$

in which ΔH_f is the heat of formation of the compound formed from the elements in their standard states, S is the sublimation energy of the metal, and I is its ionization potential. The $-\Delta H^{**}$ functions for various series of compounds of a given set of cations (or anions) yield straight lines if plotted one against the other providing certain conditions of crystal structure are met. It follows that if we were to attempt to relate the $-\Delta H^{**}$ values of a series of compounds, of which some members possess π-bonds, to those of a series in which no such additional bonding is possible, the singular cases would be revealed. Compounds of third period elements may be thus compared with the corresponding compounds of group II*a* elements, for example. In Fig. 1 the $-\Delta H^{**}$ values for third period halides are plotted against $-\Delta H^{**}$ for the corresponding oxides. Figure 2 illustrates similar relations

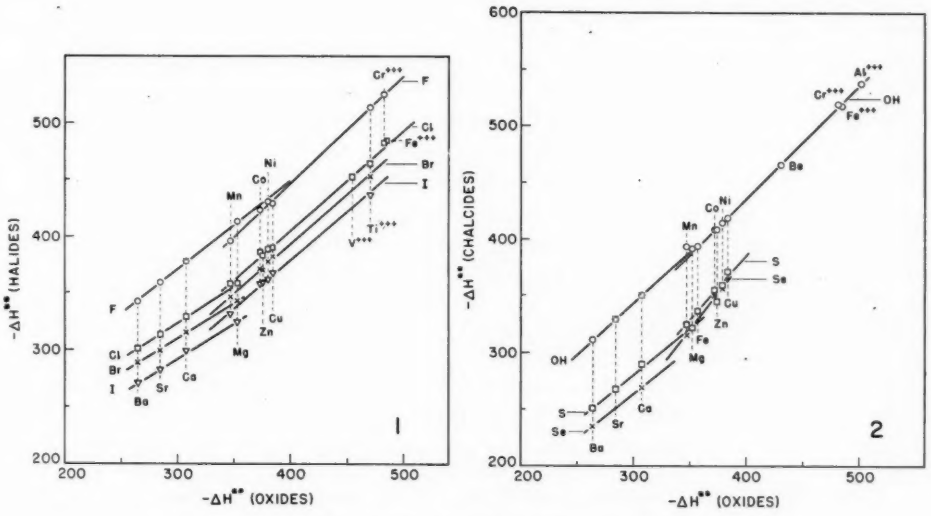

FIG. 1. Plot of $-\Delta H^{**}$ for halides against $-\Delta H^{**}$ for corresponding oxides.
FIG. 2. Plot of $-\Delta H^{**}$ for hydroxides and chalcides against $-\Delta H^{**}$ for oxides.

among the chalcides. The values for these plots are expressed in kilocalories/gram equivalent and were calculated from heats of formation published by the National Bureau of Standards, Washington (10).

An immediate conclusion which may be drawn is that there is no marked difference between the plots for fluorides and hydroxides, for which π-bonding is not possible, and the plots for anions containing vacant *d*-orbitals. The values for transition metals fall on lines which, however, are not extensions of the group II*a* plots. It must now be shown

that the deviations can be explained on the basis of a model simply involving polarization of the anion.

The value of $-\Delta H^{**}$ for nickel oxide is greater than that for magnesium oxide. This, like the higher ionization potential of nickel, may be interpreted as the result of a greater effective nuclear charge. This means that ions containing outer *d*-electrons possess a greater polarizing power than inert gas type ions of the same charge and radius. It is here assumed that the energy of polarization per gram equivalent in the halides and hydroxides is $N\alpha(Z+z)^2/r^4$, where N is the number of ion–dipole interactions per gram equivalent, α is the polarizability of the anion, Z is the charge on the cation, z is the additional effective nuclear charge due to the presence of outer *d*-electrons, and r is the sum of the cation and anion radii. This expression may be written

$$P = N\alpha[Z^2\{1+(z/Z)\}^2]/r^4. \qquad [1]$$

If z be assumed small with respect to Z, the increase in energy of polarization due to the additional effective nuclear charge is approximated by

$$\Delta P = N\alpha z^2/r^4. \qquad [2]$$

Let us now consider the compounds of cobalt, for example, and assume that the radius of the bivalent ion is 0.82 Å. The contribution to $-\Delta H^{**}$ for the oxide due to the additional polarization is $\Delta P_{CoO} = -\Delta H^{**} - (-\Delta H_h^{**})$, where $-\Delta H_h^{**}$ is the energy value for an hypothetical oxide of an inert gas type cation of radius 0.82 Å. This may be obtained from a plot of $-\Delta H^{**}$ for the group II*a* oxides against ionic radii. This procedure gives $\Delta P_{CoO} = 30.5$ kcal/equivalent. Substituting this value in the expression

$$\Delta P_{oxide} = N\alpha z^2/2r^4, \qquad [3]$$

we obtain $z = 2.46\times10^{-10}$ e.s.u. (The numeral 2 appears in the denominator of [3] because there are only half as many anions per equivalent in the oxide as in the halides.) Upon substituting this value and the appropriate values of α and r obtained from the literature (11, 12) in [2] we may calculate ΔP values for the other cobalt compounds.

Values of $-\Delta H_h^{**}$ for hypothetical halides of an inert gas type cation of radius 0.82 Å are read from plots of the values for group II*a* halides against ionic radii. When the polarization energies calculated as described above are added to these $-\Delta H_h^{**}$ values, calculated values of $-\Delta H^{**}$ are obtained. These are shown in Table I along with the

TABLE I

Comparison of calculated and experimental values (in kilocalories/gram equivalent) of $-\Delta H^{**}$ for third period elements

	Cation									
	Mn^{++}		Co^{++}		Ni^{++}		Cr^{++}		Fe^{+++}	
Anion	Calc.	Exptl.	Calc.	Exptl.	Calc.	Exptl.	Calc.	Exptl.	Calc.	Exptl.
F	406	395	428	423	433	430	524±5	525	—	—
Cl	363	358	387	385	387	388	483	482	487	484
Br	349	346	375	372	375	377	—	—	—	—
I	335	330	362	356	361	361	—	—	—	—
S	327	325	355	353	355	359	—	—	—	—

experimental values. The agreement is very good for a calculation of this type. However, it was not possible to account for the large lattice energies of selenides.*

Points for the second and third long period elements fall above the lines of Figs. 1 and 2 in those cases where π-bonding is a possibility. The heights of the points above the curves are given in Table II. In keeping with expectation for π-bonding, the energy differences

TABLE II

Experimental values of $-\Delta H^{**}$ minus the graphical interpolation values in kilocalories/gram equivalent

	Sn (II)	Pb (II)	Pd (II)	Cd	Hg (II)	In	Tl
F	—	4	—	4	—	6	—
Cl	−4	3	4	5	7	3	8
Br	−4	5	4	7	10	4	10
I	−2	8	—	10	19	5	—
OH	0	−1	0	0	—	−1	—
S	11	13	—	10	13	—	—

are greatest for third long period elements in iodides and sulphides and further, the hydroxides exhibit no enhancement of lattice energies in contrast to the sulphides. The differences for PbF_2 and CdF_2 are anomalous if we assume the origin of the effects to be in π-bonding but it should be noted that there is more uncertainty about the position of the best fluoride plot than there is about any of the others. Moreover, the differences in the case of these two compounds are equally anomalous even if a polarization argument is adduced to explain these effects.

Although it is here suggested that the plots of $-\Delta H^{**}$ provide some argument for π-bonding, it is recognized that polarization or other causes may well be responsible. It is evident, at least, that the participation of the *d*-electrons is very slight at most in these compounds.

We should expect *d*-electron donation to be more favorable when the metal atom possesses a smaller formal charge. Thermochemical data for compounds of zero-valent elements and, next best, for monovalent ones would be very welcome in this regard. At present we may make comparisons between compounds of silver and those of the alkali elements. Proceeding in this case as described above for bivalent ions, the data in Table III were obtained.

TABLE III

Comparison of calculated and experimental values of $-\Delta H^{**}$ for compounds of silver

Anion	Calc.	Exptl.
Cl	266	268.5
Br	259	269
I	248	260
S	255	253

The large discrepancies for the heavier halides favor the principle of π-bonding although this is not so for the sulphide. The explanation based upon polarization is less

*There is some evidence of participation in the bonding on the part of the d-electrons in these compounds with the NiAs structure since the magnetic moments indicate extensive superexchange among unpaired spins.

satisfactory here because the differences between calculated and experimental values are equally as great as, or greater than, those in the compounds of the neighboring ions Pd^{++} and Cd^{++}. Such a large polarizing power would be surprising in an ion which has a greater radius and is only singly charged.

We conclude that there is little or no π-bonding in the binary halides and sulphides of third period elements but there is a possible indication of such in the case of heavier elements. It is probable that extensive π-bonding occurs only in more covalent compounds in which the formal charge on the metal atom is zero or negative as is the case in the metal carbonyls.

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THE SYNTHESIS OF 3-*O*-β-D-XYLOPYRANOSYL-D-XYLOSE AND THE RECHARACTERIZATION OF SOME BENZYLIDENE DERIVATIVES OF D-XYLOSE[1]

E. J. C. CURTIS AND J. K. N. JONES

ABSTRACT

The benzylidene groups of the di-*O*,*O*-benzylidene derivative of D-xylose diethyl dithioacetal have been located at the 2,4:3,5 positions, and *not* at the 2,3:4,5 positions suggested by Zinner, Rembarz, Linke, Ulbricht. Partial hydrolysis yielded the 2,4-*O*-benzylidene compound, from which 3-*O*-β-D-xylopyranosyl-D-xylose (rhodymenabiose) was prepared.

INTRODUCTION

An attempt was made to synthesize 4-*O*-β-D-xylopyranosyl-D-xylose by condensing tri-*O*-acetyl-α-D-xylopyranosyl bromide under standard Koenigs–Knorr conditions with a D-xylose derivative which had all the hydroxyl groups substituted with the exception of that at C_4. It was also desired that this substance be in the open-chain form since experiments have indicated (2) that such reactants give greater yields than those in a ring form due to a lessening of steric hindrance in the Koenigs–Knorr reaction. It was decided that such a derivative might be synthesized using as starting material the mono-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal prepared by Zinner *et al.* (1). This was obtained in low yield by the partial hydrolysis of a di-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal formed by condensing D-xylose diethyl dithioacetal with benzaldehyde in the presence of hydrochloric acid. The above workers (1) assigned the 2,3-linkage to the mono-*O*,*O*-benzylidene compound and gave the 2,3:4,5-linkage to the di-*O*,*O*-benzylidene compound.

Accordingly, the crystalline *O*-benzylidene derivative was prepared and was acetylated to yield a crystalline di-*O*-acetyl-mono-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal, which was then demercaptalated with mercuric chloride and cadmium carbonate in dry methanol, yielding a crystalline di-*O*-acetyl-mono-*O*,*O*-benzylidene-D-xylose dimethyl acetal. Deacetylation of the latter compound produced a crystalline mono-*O*,*O*-benzylidene-D-xylose dimethyl acetal, which was then partially benzoylated using one equivalent of benzoyl chloride in pyridine at low temperature. This reaction yielded a crystalline *O*-benzoyl-*O*,*O*-benzylidene-D-xylose dimethyl acetal. From the reaction sequence, this compound should be 5-*O*-benzoyl-2,3-*O*-benzylidene-D-xylose dimethyl acetal, which has all its hydroxyl groups substituted except that at C_4, a position which is sterically unhindered owing to the absence of a hemiacetal ring. Hence, it was a suitable intermediate for the desired synthesis of 4-*O*-β-D-xylopyranosyl-D-xylose.

O-Benzoyl-*O*,*O*-benzylidene-D-xylose dimethyl acetal was therefore condensed with tri-*O*-acetyl-α-D-xylopyranosyl bromide under standard Koenigs–Knorr conditions (3). The reaction product was deacetylated and debenzoylated by warming it with sodium hydroxide solution, and benzylidene and acetal groups were subsequently removed by hydrolysis with warm dilute sulphuric acid. There resulted a syrup which, from chromatographic examination, consisted mainly of xylose and a disaccharide with an unexpectedly high mobility on chromatograms. The mixture was separated by chromatography on a cellulose column, and the disaccharide component was isolated as a syrup which crystallized. The disaccharide showed on chromatograms a single spot with an unexpectedly

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Contribution from the Department of Chemistry, Queen's University, Kingston, Ontario.

high mobility (R_{xylose} 0.53, solvent A; 0.64, solvent B), whereas the mobility of the expected 4-*O*-β-D-xylopyranosyl-D-xylose* was much lower (R_{xylose} 0.20, solvent A; 0.38, solvent B). The disaccharide was, in fact, shown to be 3-*O*-D-xylopyranosyl-D-xylose, from its oxidation with sodium periodate and the formation of an osazone. The specific rotation ($-35° \rightarrow -22°$ (constant)) showed the compound to be 3-*O*-β-D-xylopyranosyl-D-xylose. 3-*O*-α-D-Xylopyranosyl-D-xylose has a specific rotation of $+106° \rightarrow +118°$ (constant) (5).

Howard (6) has obtained a xylose-containing disaccharide by partial hydrolysis of a xylan produced by *Rhodymenia palmata*. This disaccharide, "rhodymenabiose", has the same physical constants as the synthetic xylobiose, which is now identified as 3-*O*-β-D-xylopyranosyl-D-xylose. The structure of "rhodymenabiose" was tentatively suggested to be 3-*O*-β-D-xylopyranosyl-D-xylose but no attempts were made to characterize it (6).

As the disaccharide obtained in the Koenigs–Knorr synthesis was not that expected, it was decided to re-examine the structure of the *O*-benzoyl-*O*,*O*-benzylidene-D-xylose dimethyl acetal (VII) used in the condensation. Compound VII was methylated using silver oxide and methyl iodide. The compound proved difficult to methylate, but, after three repetitions of the methylation procedure, a crystalline monomethyl ether was isolated. Debenzoylation yielded a crystalline mono-*O*,*O*-benzylidene-*O*-methyl-D-xylose dimethyl acetal which, on acid hydrolysis, was converted into crystalline 3-*O*-methyl-D-xylose (7).

Compounds VI and VII were then treated with toluene-*p*-sulphonyl chloride, and mono-*O*,*O*-benzylidene-di-*O*-tosyl-D-xylose dimethyl acetal and *O*-benzoyl-*O*,*O*-benzylidene-*O*-tosyl-D-xylose dimethyl acetal were formed, respectively. Reaction of the first of these two tosyl derivatives with sodium iodide under standard conditions (8) showed the presence of one primary hydroxyl group, while, under the same conditions, the second compound showed primary hydroxyl groups to be absent.

These data, together with the formation of 3-*O*-methyl-D-xylose, characterize VI as 2,4-*O*-benzylidene-D-xylose dimethyl acetal and VII as 5-*O*-benzoyl-2,4-*O*,*O*-benzylidene-D-xylose dimethyl acetal. Compound VI was not oxidized by sodium metaperiodate under standard conditions (9) and this is consistent with the benzylidene group in the 2,4 position, there being no free hydroxyl groups on adjacent carbon atoms in 2,4-*O*-benzylidene-D-xylose dimethyl acetal. If the benzylidene group were attached in the 2,3 position a consumption of one equivalent of periodate would be expected. Reaction of VII with tri-*O*-acetyl-α-D-xylopyranosyl bromide would therefore be expected to yield a $(1 \rightarrow 3)$-linked xylopyranosyl–xylose and not the $(1 \rightarrow 4)$-linked disaccharide originally expected in the condensation.

3-*O*-β-D-Xylopyranosyl-D-xylose was obtained in 11% yield, which is less than might be expected from a condensation involving an open-chain sugar (2). However, although the sugar was in the open-chain form, the hydroxyl group involved in the reaction—that at C_3—was part of a ring system which involved the benzylidene group, and, thus, reaction was sterically hindered to some extent. That the hydroxyl group at C_3 was not very reactive was shown by the difficulties encountered in methylating and tosylating this group.

In order to determine if the mono-*O*,*O*-benzylidene derivatives III, IV, and V had the benzylidene group in the 2,4 position or if the group had wandered during one of the interconversions, the structure of mono-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal

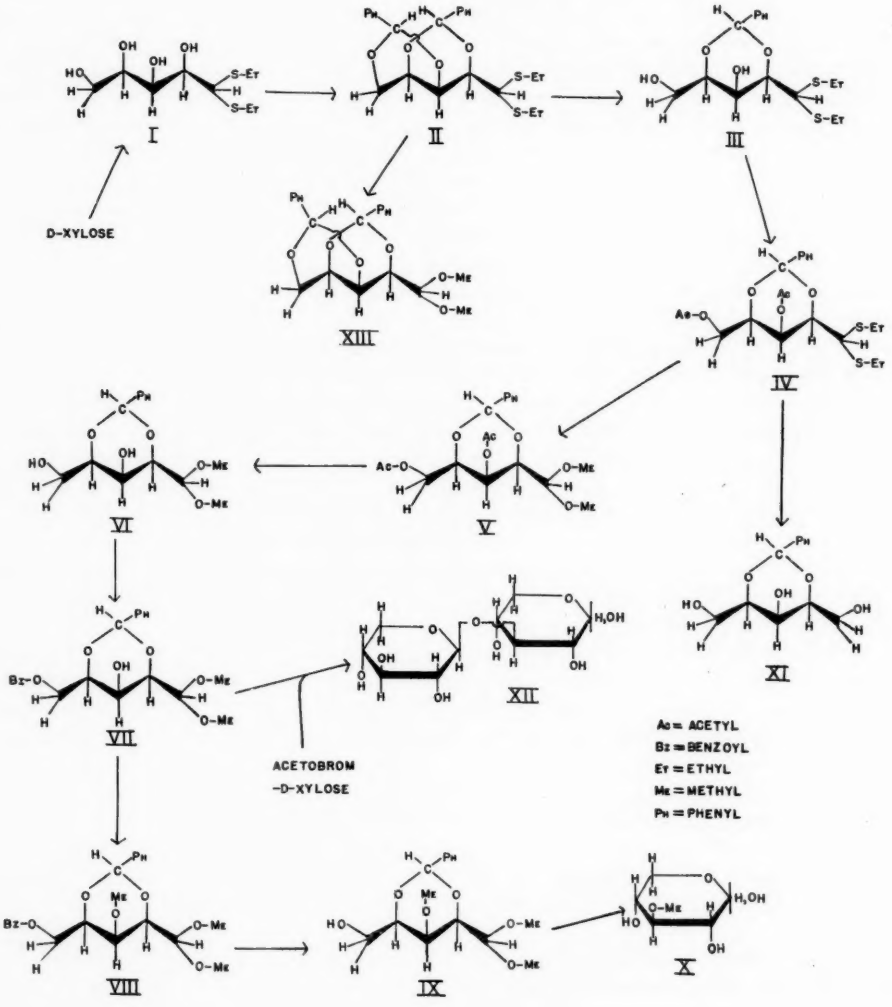
**Prepared by partial hydrolysis of a hemicellulose from aspen (Populus tremuloides) (4).*

(III) was examined. The formulation of this as the 2,3-*O*-benzylidene isomer (Zinner *et al.* (1)) rests upon its oxidation with lead tetraacetate, when one equivalent of the oxidizing agent was consumed and formaldehyde (amount not stated) was liberated. This consumption of oxidant was interpreted to mean that the compound contained hydroxyl groups on adjacent carbon atoms. It is known, however (10, 11), that lead tetraacetate oxidation of thioacetals is an unreliable guide to the presence of hydroxyl groups on adjacent carbon atoms, due to the oxidation of the dithioacetal grouping. In particular, Huebner *et al.* (10) have examined the lead tetraacetate oxidation of the analogous mono-*O*,*O*-benzylidene derivatives of L-arabinose diethyl dithioacetal. With the fully substituted di-*O*-benzylidene-L-arabinose diethyl dithioacetal, one equivalent of oxidant was consumed rapidly (5 minutes) and a further one equivalent was consumed slowly (2 hours). This consumption they assumed to be due to oxidation of the dithioacetal group—a rapid oxidation of one sulphur atom followed by a slower oxidation of the second. With mono-*O*,*O*-benzylidene-L-arabinose diethyl dithioacetal, a consumption of two equivalents (rapid) plus two equivalents (slow) of lead tetraacetate per mole was observed, which, it was assumed, was due partially to oxidation of the thioacetal groups (one fast plus one slow) and partially to oxidation of hydroxyl groups on adjacent carbon atoms (also, one fast plus one slow). It would, therefore, seem from these observations that the data reported (1) for the lead tetraacetate oxidation of mono-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal would support the view that this compound does *not* possess hydroxyl groups on adjacent carbon atoms, rather than the reverse.

By demercaptalation of II, Zinner *et al.* (1) obtained a 2,3,4,5-di-*O*-benzylidene-*aldehydo*-D-xylose which, by heating with methanol, they converted into a crystalline 2,3,4,5-di-*O*-benzylidene-D-xylose methyl hemiacetal. They point out that the physical constants (melting point and specific rotation) of this compound are very similar to those of a 2,3,4,5-di-*O*-benzylidene-L-xylose methyl hemiacetal, which was prepared by lead tetraacetate oxidation of a 1,2,3,4-di-*O*-benzylidene-D-glucitol and subsequent treatment with methanol (Wolfe, Hann, and Hudson (12)). However, although the structures of these L-xylose and D-glucitol derivatives were at first in doubt, di-*O*,*O*-benzylidene-D-glucitol was later shown by Angyal and Lawler (13) to have the benzylidene groups spanning the 1,3 and 2,4 hydroxyl groups. These workers isolated, by partial hydrolysis, 2,4-*O*-benzylidene-D-glucitol, the structure of which was previously known. It therefore follows that the 2,3,4,5-di-*O*-benzylidene-L-xylose methyl hemiacetal must, in fact, have the 2,4:3,5 structure, and the identity of physical constants of this and di-*O*,*O*-benzylidene-D-xylose methyl hemiacetal would indicate that the latter compound *too* has the 2,4:3,5 structure.

Finally, to characterize compounds II, III, and IV, di-*O*-acetyl-*O*,*O*-benzylidene-D-xylose diethyl dithioacetal (IV) was demercaptalated, using mercuric chloride and cadmium carbonate in aqueous acetone (14), to yield di-*O*-acetyl-*O*,*O*-benzylidene-*aldehydo*-D-xylose. This compound was reduced with sodium borohydride in aqueous methanol, and the resulting alkaline solution was heated on a water bath to complete removal of the acetyl groups. The solution was deionized and evaporated to a crystalline residue which was shown to be the well-characterized, symmetrical 2,4-*O*-benzylidene-xylitol (XI) (15, 16). Compound IV is, therefore, 3,5-di-*O*-acetyl-2,4-*O*-benzylidene-D-xylose diethyl dithioacetal, whereas compounds II and III must be 2,4:3,5-di-*O*-benzylidene-D-xylose diethyl dithioacetal and 2,4-*O*-benzylidene-D-xylose diethyl dithioacetal, respectively.

The real sequence of reactions in the synthesis of the disaccharide was therefore:



Demercaptalation of 2,4:3,5-di-*O*-benzylidene-D-xylose diethyl dithioacetal (II), using mercuric chloride and cadmium carbonate in dry methanol, yielded crystalline 2,4:3,5-di-*O*-benzylidene-D-xylose dimethyl acetal (XIII). This was shown to be identical with the di-*O*,*O*-benzylidene-D-xylose dimethyl acetal prepared by Breddy and Jones (17) by condensing D-xylose with benzaldehyde and methanol in the presence of hydrochloric acid, and used by them for the estimation of xylose. The structure of this compound was not previously known.

Zinner *et al.* (1), in addition to preparing the mono-*O*,*O*-benzylidene (III) and di-*O*,*O*-benzylidene (II) derivatives of D-xylose diethyl dithioacetal, also prepared, in the same

way, a mono-*O,O*-benzylidene-D-xylose dipropyl dithioacetal, and di-*O,O*-benzylidene derivatives of the methyl, propyl, *n*-butyl, isobutyl, and benzyl dithioacetals of D-xylose. These, like II and III, were assigned the 2,3-linkage (in the case of the mono-*O,O*-benzylidene-D-xylose dipropyl dithioacetal) and the 2,3:4,5-linkage (in the case of the di-*O,O*-benzylidene derivatives of the various D-xylose dithioacetals). No attempt has been made to investigate the structures of these derivatives, but it would seem likely that they, like II and III, have the di-*O,O*-benzylidene groups in the 2,4:3,5 positions and the mono-*O,O*-benzylidene group in the 2,4 position.

EXPERIMENTAL

Optical rotations were measured at 21°± 2° C and in water unless otherwise stated. Solutions were concentrated under reduced pressure at 40° C. The following solvent systems (v/v) were used to separate sugars on paper chromatograms: (A) 1-butanol:ethanol:water, 3:1:1; (B) ethyl acetate:acetic acid:water, 9:2:2. Sugars were detected with the *p*-anisidine hydrochloride spray (18). Infrared absorptions were measured either as solutions in chloroform or as a powder (mulled in a potassium bromide pellet) on a Perkin–Elmer Model 21 spectrophotometer.

D-*Xylose Diethyl Dithioacetal* (I)

D-Xylose (100 g), ice-cold ethanethiol (100 g), and ice-cold hydrochloric acid (100 ml) were shaken together in a stoppered flask and the resulting mixture was left overnight at room temperature. The solution was diluted with water (200 ml) and then deionized by passage through Amberlite IR 400 (OH) resin. The eluate was concentrated to a syrup which was diluted with ether (200 ml) and refrigerated. The product crystallized and was filtered and washed with ether. Yield: 140 g (82%), m.p. 63–64° C.

2,4:3,5-Di-O-benzylidene-D-*xylose Diethyl Dithioacetal* (II)

D-Xylose diethyl dithioacetal (50 g) was dissolved in redistilled benzaldehyde (500 ml) with warming and after being cooled to 5° C in ice, dry hydrogen chloride was bubbled through the solution. After 5 minutes a semisolid mass was formed which was diluted with ethanol (400 ml), shaken, and filtered under suction. The crude crystalline product was dried *in vacuo* and recrystallized from boiling ethanol. Yield: 40.6 g (48%), m.p. 181–182° C, $[\alpha]_D$ −36° (*c*, 4.95, chloroform). Literature (1) gives a melting point of 181–182° C and $[\alpha]_D$ −36.7° (*c*, 2.53, chloroform).

2,4-O-Benzylidene-D-*xylose Diethyl Dithioacetal* (III)

The conditions given by Zinner *et al.* (1) (hydrolysis with ethanol (3 parts) plus 80% acetic acid (4 parts) for 1 hour) were found to produce only a small yield of mono-*O,O*-benzylidene compound. After various hydrolyzing mixtures were tried, the following conditions were found to be most satisfactory. A mixture of di-*O,O*-benzylidene-D-xylose diethyl dithioacetal (80 g), acetone (1800 ml), and 80% acetic acid (1200 ml) was refluxed for 3 hours. The resulting solution was cooled to 10° C in ice water, and unhydrolyzed di-*O*-benzylidene material was removed by filtration. The filtrate was concentrated to a small volume and the acid was neutralized with cold potassium hydroxide solution followed by sodium bicarbonate solution. The aqueous solution was extracted with chloroform (3×100 ml) and the combined extracts were washed once with water, dried ($MgSO_4$), and filtered. Evaporation of the filtrate yielded a crystalline residue which was extracted with boiling water. The aqueous extract was cooled and mono-*O,O*-benzylidene-D-xylose diethyl dithioacetal was removed by filtration. Yield: 5.5 g (43% after recovery

of unchanged di-*O*-benzylidene compound). After recrystallization from water and from ether it had a melting point of 146–147° C and $[\alpha]_D$ −22° (*c*, 2.0, methanol). Literature (1) gives a melting point of 147° C and $[\alpha]_D$ −23.3° (*c*, 2.76, methanol).

3,5-Di-O-acetyl-2,4-O-benzylidene-D-*xylose Diethyl Dithioacetal* (IV)

2,4-*O*-Benzylidene-D-xylose diethyl dithioacetal (29.4 g) was dissolved in dry pyridine (300 ml), and acetic anhydride (250 ml) was added dropwise to the solution while this was cooled and stirred mechanically. After the addition (1 hour) stirring was continued for 1 hour more and the solution was then left at room temperature (16 hours), after which it was poured with stirring into ice water (1200 ml). The product precipitated from solution and was collected, washed with water, and air-dried. After recrystallization from methanol it weighed 35 g (96%) and had a melting point of 92.5–93° C and $[\alpha]_D$ −2° (*c*, 4.35, chloroform). Anal. Calc. for $C_{20}H_{28}O_6S_2$: S, 14.95. Found: S, 14.96.

3,5-Di-O-acetyl-2,4-O-benzylidene-D-*xylose Dimethyl Acetal* (V)

3,5-Di-*O*-acetyl-2,4-*O*-benzylidene-D-xylose diethyl dithioacetal (10 g) was dissolved in dry methanol (300 ml). Cadmium carbonate (90 g) was added to the solution, which was stirred vigorously, and a solution of mercuric chloride (90 g) in dry methanol (100 ml) was added portionwise at room temperature. After the addition (1 hour) stirring was continued for a further 15 hours at room temperature and the mixture was then refluxed (40 minutes). The solution was filtered through a Celite pad and the precipitate was washed with methanol. Filtrate and washings were concentrated to ca. 200 ml and this solution was poured into 500 ml each of chloroform and water. The chloroform layer was extracted with water until chloride-free, dried ($MgSO_4$), filtered, and evaporated to a syrup (9.6 g). This crystallized when allowed to stand in the desiccator. The crude product was recrystallized from methanol. Yield: 6.5 g (76%), m.p. 82.5–83.5° C, $[\alpha]_D$ +34° (*c*, 4.3, chloroform). Anal. Calc. for $C_{18}H_{24}O_8$: C, 58.71; H, 6.52; S, nil. Found: C, 58.94; H, 6.34; S, 0.07.

2,4-O-Benzylidene-D-*xylose Dimethyl Acetal* (VI)

3,5-Di-*O*-acetyl-2,4-*O*-benzylidene-D-xylose dimethyl acetal (1.45 g) was dissolved in dry methanol (30 ml) and the resulting solution was cooled to 5° C. Methanolic sodium methoxide (0.88 *N*, 1 ml) was added and the solution was then stored at 7° C (22 hours) in a stoppered flask. The solution was neutralized (glacial acetic acid and then sodium bicarbonate solution) and then evaporated to dryness, to yield a white crystalline residue. This was recrystallized from ether. Yield: 0.90 g (81%), m.p. 142–143° C, $[\alpha]_D$ +32° (*c*, 1, methanol). Infrared examination showed no absorption for carbonyl group (1700–1750 cm^{-1}). Anal. Calc. for $C_{14}H_{20}O_6$: C, 59.16; H, 7.04. Found: C, 59.41; H, 7.15.

5-O-Benzoyl-2,4-O-benzylidene-D-*xylose Dimethyl Acetal* (VII)

The method used was similar to that described by Zinner *et al.* (19). 2,4-*O*-Benzylidene-D-xylose dimethyl acetal (4.8 g) was dissolved in dry pyridine and the solution was cooled to −15° C and stirred mechanically. A solution of redistilled benzoyl chloride (2.5 g) in dry pyridine (10 ml) was then added dropwise while the temperature was kept below −10° C (40 minutes). The solution was stored at −5° C (4 hours), then at 0° C (16 hours), and finally at room temperature (2 hours), after which the product was precipitated by pouring the solution into ice water (700 ml). The mixture was left at room temperature (2 hours) to complete hydrolysis of excess benzoyl chloride and was then filtered. The precipitate was washed with water and after drying *in vacuo* it weighed 4.5 g (69%). The material was recrystallized from methanol and then had a melting point of 120–121° C

and $[\alpha]_D$ +44° (*c*, 4.06, chloroform). Anal. Calc. for $C_{21}H_{24}O_7$: C, 64.94; H, 6.19. Found: C, 65.21; H, 6.40.

*Condensation of 5-O-Benzoyl-2,4-O-benzylidene-*D*-xylose Dimethyl Acetal with Tri-O-acetyl-α-*D*-xylopyranosyl Bromide*

5-*O*-Benzoyl-2,4-*O*-benzylidene-D-xylose dimethyl acetal (8 g) was dissolved in dry, alcohol-free chloroform (80 ml) and to the solution were added Drierite (10 g) and silver oxide (10 g). The solution was shaken in the dark to dry the reagents. After 1 hour tri-*O*-acetyl-α-D-xylopyranosyl bromide (20) (14.4 g) in dry, alcohol-free chloroform (50 ml) and a few crystals of iodine were added and the shaking was continued at room temperature. After 48 hours a test for ionizable bromine was negative. After a further 48 hours the solution was filtered through a Celite pad, and evaporated to a syrup which was dissolved in a solution of water (4 parts) plus methanol (1 part) (200 ml) containing sodium hydroxide (8 g). The solution was warmed on a water bath at 70° C (4 hours), after which it was cooled and deionized by passage through columns of Amberlite IR 400 (OH) and IRC 50 (H) resins in series; the former resin also removed the bulk of the free xylose. The neutral eluate was concentrated to a syrup (5.4 g) which was dissolved in aqueous *N*/100 sulphuric acid, and the resulting solution was warmed at 70° C (5 hours). The solution was cooled and deionized by passage through Duolite A-4 (OH) resin. Concentration of the eluate yielded a syrup which was shown by paper chromatography (solvent B) to consist, in addition to xylose, of a main component with R_{xylose} 0.64 and smaller components with R_{xylose} 1.23 and 0.38.

The syrup was partially fractionated by chromatography on a column of Dowex 50W resin (94×3.3 cm) (21) and fractionation and concentration of the eluate yielded a syrup (2.3 g) which was enriched with the chromatographically slower-moving components. Part of this (0.9 g) was fractionated by chromatography on a cellulose column (40×3 cm) (22) using as eluent 1-butanol half saturated with water.

Fraction 1: A syrup (0.62 g) which chromatographic examination showed to consist mainly of xylose, plus a much smaller component with R_{xylose} 1.23 (solvent B). This latter was presumed to be a partially unhydrolyzed xylose acetal or ester.

Fraction 2: A syrup (0.255 g) which crystallized on standing. Chromatography showed a single spot with R_{xylose} 0.53 (solvent A) and 0.64 (solvent B).

Fraction 3: A syrup (0.04 g) which chromatographic examination showed to consist of a major component with R_{xylose} 0.26 (solvent A) and 0.38 (solvent B) and a minor component with R_{xylose} 0.53 (solvent A) and 0.64 (solvent B). This fraction was not examined further.

Fraction 2 was recrystallized from methanol and then had a melting point of 192–193° C and $[\alpha]_D$ −35° (4 minutes) → −22° (constant) (*c*, 2.5).

Acid hydrolysis (*N* sulphuric acid at 80° C for 12 hours) showed only xylose to be present. Anal. Calc. for $C_{10}H_{18}O_9$: C, 42.53; H, 6.43. Found: C, 42.71; H, 6.56.

Periodate Oxidation of Fraction 2 (XII)

Portions (ca. 10 mg) of fraction 2 were oxidized with aqueous sodium metaperiodate using standard procedures, (*a*) in acid pH unbuffered conditions, and (*b*) in solution buffered at pH 8. The periodate uptake was estimated by the method of Neumüller and Vasseur (9) and the formic acid liberation by that of Andrews *et al.* (23).

The molar periodate consumptions per mole of disaccharide were as follows: (*a*) in unbuffered solution, 1.38 (after 14 minutes), 1.75 (39 minutes), 2.36 (1.2 hours), 2.41 (2 hours), 2.78 (4 hours), 2.97 (7.5 hours), and 3.04 (21 hours); (*b*) in solution buffered at

pH 8, 1.48 (after 13 minutes), 2.88 (38 minutes), 7.11 (4 hours), and 8.73 (21 hours). The molar production of formic acid in unbuffered solution was 0.47 (after 39 minutes), 0.58 (1.2 hours), 0.82 (2 hours), 0.99 (4 hours), 1.20 (7.5 hours), and 1.33 (21 hours). Extrapolation of the flat part of the curve to zero time gave values for the experiment in unbuffered solution: uptake of metaperiodate, 2.95 moles per mole of sugar; formic acid production, 1.05 moles per mole of sugar.

The theoretical results for oxidation of a xylobiose in acid pH unbuffered conditions are: for (1 → 2) and (1 → 3)-linked disaccharides, 3 moles of periodate consumed per mole of sugar and 1 mole of formic acid liberated per mole of sugar, and for (1 → 4)- and (1 → 5)-linked disaccharides, 4 moles of periodate per mole and 2 moles of formic acid per mole assuming hydrolysis of formyl ester is slow. The extrapolated values obtained therefore indicated that the glycosidic linkage was either (1 → 2) or (1 → 3), and the overoxidation values showed that it was in fact the latter. There was very little overoxidation in unbuffered solution (only 3.05 moles of periodate were consumed after 21 hours) and this is typical of (1 → 3)-linked disaccharides where initial oxidation produces a formyl ester-containing residue which is stable under the acid pH which results in unbuffered conditions. With (1 → 3)-linked sugars overoxidation is rapid in alkaline solution as the formyl ester is unstable under these conditions, and fraction 2 showed a rapid overoxidation in solution buffered at pH 8. It is known (25) that with (1 → 2)-linked disaccharides overoxidation is rapid even in acid conditions because the residue resulting from initial oxidation does not contain a formyl ester and hence is unstable.

*3-O-β-*D*-Xylopyranosyl-*D*-xylose Phenyl Osazone*

The sugar (50 mg) was dissolved in water (0.5 ml) and acetic acid (0.2 ml) plus phenyl hydrazine (0.15 ml) were added. The solution was warmed at 70° C (1.5 hours) and then cooled and diluted with water (3 ml). An oil separated which crystallized on standing. This was recrystallized from hot water and then had a melting point of 194–196° C and $[\alpha]_D$ +47° (*c*, 1.25, pyridine). Anal. Calc. for $C_{22}H_{28}O_7N_4$: N, 12.18. Found: N, 12.28.

*5-O-Benzoyl-2,4-O-benzylidene-3-O-methyl-*D*-xylose Dimethyl Acetal* (VIII)

5-*O*-Benzoyl-2,4-*O*-benzylidene-D-xylose dimethyl acetal (1.19 g) was dissolved in methyl iodide (25 ml). Silver oxide (2 g), Drierite (2 g), and a few crystals of iodine were added to the solution, which was then shaken in the dark (21 hours) and finally boiled under reflux (5 hours) with exclusion of moisture. The solution was filtered and the residue was washed with acetone. Filtrate and washings were evaporated to dryness, yielding a partially crystalline residue which was finally dried *in vacuo*. It then weighed 1.17 g and, on infrared analysis, showed appreciable amounts of hydroxyl group to be still present. The above methylation procedure was repeated twice (the materials were boiled under reflux for a total time of 70 hours), after which a crystalline product (1.17 g, total yield i.e. recovery, 96%) was obtained which, when subjected to infrared analysis, showed only a small peak corresponding to free hydroxyl group. The product was dissolved in warm ethanol and cooling the resulting solution induced partial crystallization. The crystalline material was filtered, washed with ethanol, and dried. It weighed 0.37 g and had a melting point of 137–138° C. Further cooling of the solution produced a second crop of crystals (0.16 g, m.p. 137–138° C), after which a third crop was obtained which had a melting point of 90–98° C and was presumably a mixture of methylated and nonmethylated materials. The material with a melting point of 137–138° C was combined (0.53 g, 43%) and recrystallized from ethanol. It then had a melting point of 138–139° C and $[\alpha]_D$ +40° (*c*, 2.5, chloroform). Infrared examination showed there to be no peak for

free hydroxyl group present. Anal. Calc. for $C_{22}H_{26}O_7$: OMe, 23.14. Found: OMe, 23.32.

*2,4-O-Benzylidene-3-O-methyl-*D*-xylose Dimethyl Acetal* (IX)

5-*O*-Benzoyl-2,4-*O*-benzylidene-3-*O*-methyl-D-xylose dimethyl acetal (0.34 g) was dissolved in a solution of ethanol (10 ml) plus aqueous *N*/10 sodium hydroxide solution (10 ml) and the materials were then warmed on a water bath at 70° C (2 hours). The solution was cooled and concentrated to dryness and yielded a residue which was dissolved in chloroform (20 ml). The resulting solution was washed twice with water, dried ($MgSO_4$), filtered, and evaporated to dryness. A crystalline residue was obtained (0.22 g, 88%) which after recrystallization from ether had a melting point of 134–135° C and $[\alpha]_D$ +9° (*c*, 1.55, methanol). Anal. Calc. for $C_{15}H_{22}O_6$: OMe, 31.2. Found: OMe, 31.0.

*3-O-Methyl-*D*-xylose* (X)

2,4-*O*-Benzylidene-3-*O*-methyl-D-xylose dimethyl acetal (0.22 g) was added to aqueous *N*/60 sulphuric acid (10 ml) and sufficient methanol was then added to dissolve the material. The solution was warmed at 70° C (5 hours), cooled, and deionized by passage through Duolite A-4 resin (OH form). Concentration of the eluate yielded a colorless syrup (0.12 g 95%), which crystallized completely after nucleation with authentic 3-*O*-methyl-D-xylose. After recrystallization from acetone the material had a melting point of 96° C and mixed melting point with authentic 3-*O*-methyl-D-xylose 95–96° C, $[\alpha]_D$ +57° (4 minutes) → +19° (constant) (*c*, 1.6). Literature gives a melting point of 103–104° C and $[\alpha]_D$ +52.2° → +14.8° (constant) (7), m.p. 95° C, and $[\alpha]_D$ +45 °→ +19° (constant) (26). The infrared spectra of this material and of authentic 3-*O*-methyl-D-xylose were identical.

*2,4-O-Benzylidene-3,5-di-O-tosyl-*D*-xylose Dimethyl Acetal*

2,4-*O*-Benzylidene-D-xylose dimethyl acetal (1.07 g) was dissolved in dry pyridine (4 ml) and the solution was cooled to 10° C. A solution of toluene-*p*-sulphonyl chloride (1.6 g) in pyridine (3 ml) was then added and the material was left in a stoppered flask at room temperature (2 hours). Excess tosyl chloride was hydrolyzed by the slow dropwise addition of water, and the mixture was then poured into an excess (100 ml) of water. The resulting solution was extracted with chloroform and the combined chloroform extracts were washed successively with solutions of dilute sulphuric acid and sodium bicarbonate, and finally with water. The solution was dried, filtered; and evaporated to a hard glass (1.97 g). Infrared examination of this showed a slight peak for free hydroxyl group so the tosylation procedure was repeated; a colorless glass (1.93 g, 86%) was thus obtained. This showed no peak corresponding to free hydroxyl group on infrared examination, $[\alpha]_D$ +14° (*c*, 3.4, chloroform). Anal. Calc. for $C_{28}H_{32}O_{10}S_2$: S, 10.77. Found: S, 10.80.

*Reaction of 2,4-O-Benzylidene-3,5-di-O-tosyl-*D*-xylose Dimethyl Acetal with Sodium Iodide*

The conditions used were those recommended by Tipson (8). 2,4-*O*-Benzylidene-3,5-di-*O*-tosyl-D-xylose dimethyl acetal (0.23 g) plus a 10% solution of sodium iodide in acetone (2.3 ml) were sealed in a glass tube, which was then placed in a boiling water bath (2 hours). After being cooled to room temperature the materials were filtered and the precipitated sodium toluene-*p*-sulphonate was washed with acetone. Yield: 0.03 g, which, allowing for the solubility of sodium toluene-*p*-sulphonate in acetone (8), represents 44% reaction from one tosyl group.

*5-O-Benzoyl-2,4-O-benzylidene-3-O-tosyl-*D*-xylose Dimethyl Acetal*

5-*O*-Benzoyl-2,4-*O*-benzylidene-D-xylose dimethyl acetal (0.30 g) was treated with

tosyl chloride using the procedure described above. After three repetitions of the tosylation process a colorless glass (0.41 g, 97%) was obtained which showed no peak corresponding to free hydroxyl group on infrared examination. It had $[\alpha]_D$ +4° (*c*, 4.62, chloroform). Anal. Calc. for $C_{28}H_{30}O_9S$: S, 5.90. Found: S, 5.88.

The material was heated with sodium iodide in acetone as described for 2,4-*O*-benzylidene-di-*O*-tosyl-D-xylose dimethyl acetal. After 2 hours in the boiling water bath no sodium toluene-*p*-sulphonate was produced.

2,4-O-Benzylidene-xylitol (XI)

3,5-Di-*O*-acetyl-2,4-*O*-benzylidene-D-xylose diethyl dithioacetal (IV) (5.15 g) was dissolved in a solution of acetone (4 parts) plus water (1 part) (80 ml). Cadmium carbonate (12 g) was added to the solution, which was stirred vigorously while a solution of mercuric chloride (12.3 g) in acetone (40 ml) was added portionwise at room temperature. After the addition (3 hours) stirring was continued for a further 15 hours at room temperature and the mixture was then boiled under reflux (40 minutes). The product was worked up as described for 2,4-*O*-benzylidene-3,5-di-*O*-acetyl-D-xylose dimethyl acetal, a clear syrup (3.7 g) being obtained.

The above crude 3,5-di-*O*-acetyl-2,4-*O*-benzylidene-aldehydo-D-xylose (3.0 g) was dissolved in 50% aqueous methanol (100 ml) and to the cooled solution was added an excess (2.2 g) of sodium borohydride. The solution was left at room temperature (3 hours) after which excess sodium borohydride was destroyed with glacial acetic acid. The solution was deionized by passage through a column of Amberlite IR-400 (OH) resin and the acidic eluate was shaken with Amberlite IR-120 (H) resin in the cold. The neutral solution was evaporated to dryness and boric acid was removed as methyl borate by codistillation with methanol (4×40 ml). A residue remained (1.05 g, 46% over-all yield) which crystallized on standing and was recrystallized from water. It had a melting point of 142–143° C and showed no optical rotation. Literature (16) gives a melting point of 143–144° C. The material formed a crystalline triacetate which had a melting point of 94–95° C and showed no optical rotation. Literature (16) gives a melting point of 94–95° C.

*2,4:3,5-Di-O-benzylidene-*D*-xylose Dimethyl Acetal* (XIII)

2,4:3,5-Di-*O*-benzylidene-D-xylose diethyl dithioacetal (5 g) was demercaptalated using mercuric chloride and cadmium carbonate in dry methanol as described for 3,5-di-*O*-acetyl-2,4-*O*-benzylidene-D-xylose diethyl dithioacetal (IV). The product was extracted in the same way yielding a white crystalline residue (3.1 g, 72%) which was recrystallized from chloroform/light petroleum. It had a melting point of 210–211° C and $[\alpha]_D$ −9° (*c*, 2.2, chloroform).

Di-*O*-benzylidene-D-xylose dimethyl acetal was also prepared as described by Breddy and Jones (17) by allowing D-xylose to react with methanol and benzaldehyde in the presence of hydrochloric acid. The white crystalline product was recrystallized from chloroform/light petroleum and had a melting point of 210–211° C undepressed by admixture with 2,4:3,5-di-*O*-benzylidene-D-xylose dimethyl acetal and $[\alpha]_D$ −9° (*c*, 2.25, chloroform). Literature (17) gives a melting point of 211° C and $[\alpha]_D$ −9° (*c*, 1.2, chloroform). The infrared spectrum of this material and that of 2,4:3,5-di-*O*-benzylidene-D-xylose dimethyl acetal were identical.

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STEREOCHEMICAL STUDIES

I. CONFIGURATIONS OF SOME 4-AZACHOLESTANES[1]

J. T. Edward and Peter F. Morand[2]

ABSTRACT

The Beckmann rearrangement of the oxime of A-nor-5β-cholestan-3-one gave 4-aza-5β-cholestan-3-one; the isomeric lactam obtained by the reduction of the oxime of 5-oxo-3,5-seco-A-norcholestan-3-oic acid with sodium in alcohol must consequently be 4-aza-5α-cholestan-3-one. Other methods of reducing the oximino acid afforded hydroxamic acids. 4-Aza-5α- and 4-aza-5β-cholestane were obtained by reduction of the appropriate lactam or hydroxamic acid.

For stereochemical studies we required 4-aza-5α-cholestane (VIII, R = H) and 4-aza-5β-cholestane (III, R = H), and hence the 5α- (VII) and 5β-lactam (II).

Bolt (1) obtained a lactam of this structure, but of undetermined configuration at the 5-position, by reduction of the oxime (V) of Windaus' keto acid (IV) (2) with sodium in ethanol. We have repeated Bolt's work and have obtained a lactam having properties agreeing closely with those reported. This lactam would be expected, from previous experience of reductions by sodium in alcohol (3, 4), to have a trans fusion of rings A and B, as shown in the structure VII. This stereochemistry is also suggested by the direct preparation of the lactam from the keto acid (IV) by the Leuckart reaction, which is known to be sensitive to steric effects (5, 6) and which with 5α-cholestan-3-one gives β-amino derivatives (7).

The structure (VII) for Bolt's lactam has now been proved by preparing the isomeric cis-fused lactam (II) by the Beckmann rearrangement of the oxime (I) of Windaus' norketone (8, 9). The configuration of the migrating center is known to be retained in the Beckmann rearrangement (10), and the migration of the 5- rather than 3-carbon atom would be expected from analogy with a similar reaction in the cholanic acid series (11). Evidence confirming this expectation is given below.

The lactams (VII and II) afforded 4-aza-5α- (VIII, R = H) and 4-aza-5β-cholestane (III, R = H) respectively on reduction with lithium aluminum hydride. The latter azasteroid could be obtained only as an oil but gave a crystalline perchlorate.

Bolt (12) claimed that hydrogenation over platinum of the oximino acid (V) gave the same lactam (VII) as did reduction with sodium in ethanol (1), together with a smaller amount of an amino acid. In our hands catalytic reduction has given as sole product a hydroxamic acid, shown to have the cis-fused structure (VI, R = H) by its further reduction with tin in hydrochloric acid to 4-aza-5β-cholestane (III, R = H). This reaction establishes also the 4-aza structure for the lactam (II) obtained by the Beckmann rearrangement. The isomeric trans-fused hydroxamic acid (IX, R = H) was obtained by reduction of the oximino acid (V) with zinc in acetic acid; its structure was established by its conversion to 4-aza-5α-cholestane (VIII, R = H) with lithium aluminum hydride. The differing stereochemical effects of the reducing agents in giving the isomeric hydroxamic acids are as expected (3).

While the infrared absorption frequencies of the carbonyl group in hydroxamic acids (13, 14) are generally lower than in the corresponding amides, the differences are considerably greater for the six-membered (VI, R = H, and IX) than for five-membered

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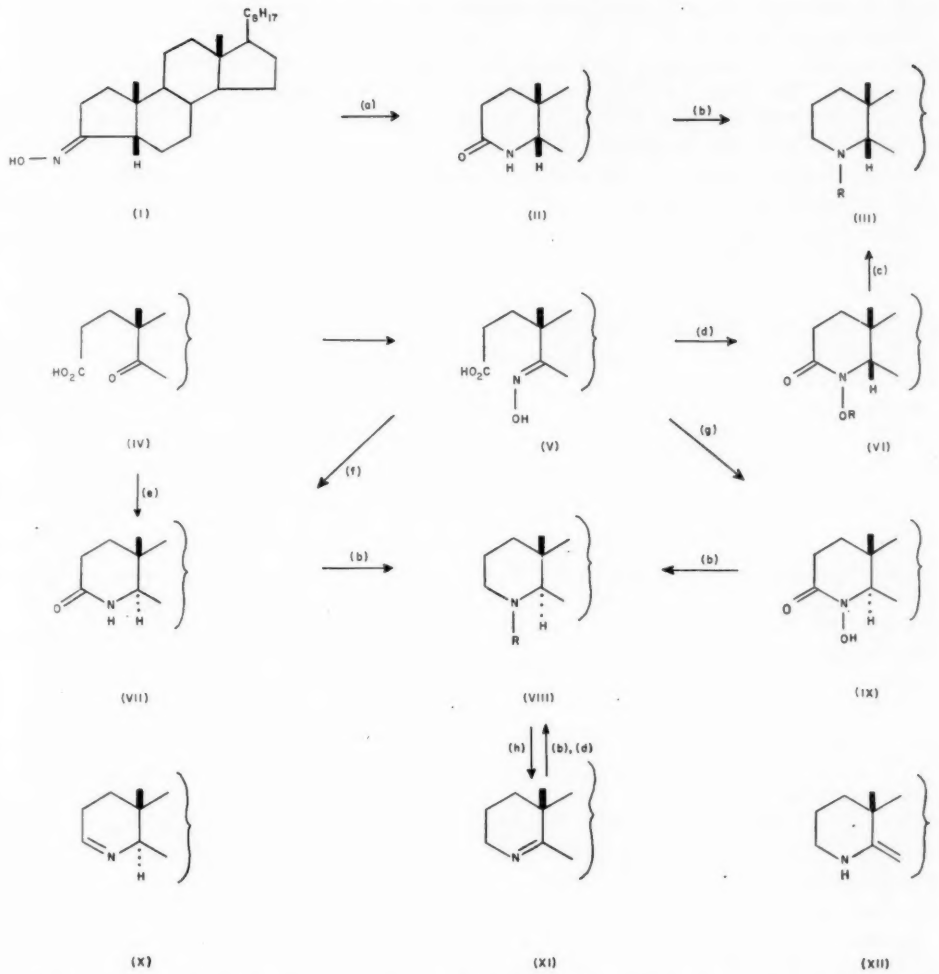


REAGENTS: (*a*) *p*-Acetamidobenzenesulphonyl chloride, (*b*) $LiAlH_4$, (*c*) Sn in HCl, (*d*) H_2/Pt, (*e*) HCO_2NH_4, (*f*) Na in ethanol, (*g*) Zn in AcOH, (*h*) $Hg(OAc)_2$.

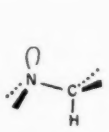
cyclic compounds (15). The very low values for the former compounds (1623 and 1630 cm^{-1}) may be due to the more effective intramolecular hydrogen bonding possible between the carbonyl and hydroxyl groups. Some support for this view comes from the higher frequencies of the derivatives (VI, R = PhCO (1677 cm^{-1}) and VI, R = Me, (1658 cm^{-1})).

Oxidation of 4-aza-5α-cholestane (VIII, R = H) with mercuric acetate afforded 4-azacholest-4-ene (XI). The alternative structures X and XII, considered less likely from the course of oxidation of analogous compounds (16), were eliminated by the absence of absorption bands in the regions 3030–3040 cm^{-1} and 800–970 cm^{-1}, characteristic of =C—H stretching vibrations and out-of-plane deformation vibrations (17),

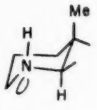
and by the absence of an absorption maximum which could be assigned to an N—H stretching vibration.

Reduction of 4-azacholest-4-ene with lithium aluminum hydride gave back 4-aza-5α-cholestane. The same product was obtained by catalytic hydrogenation which might, by analogy with the hydrogenation of cholest-4-ene in neutral solution (8), have been expected to give the 5β compound.

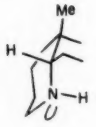
From a study of the infrared spectra of about 50 nitrogenous bases, Bohlmann (18) has suggested that the structural unit XIII, in which the tertiary hydrogen and the orbital containing the unshared electrons of the nitrogen are coplanar and trans, gives rise to an absorption band in the frequency region 2700–2800 cm^{-1}. The absence of absorption in this region by 4-aza-5α- and 4-aza-5β-cholestane may be interpreted as evidence for the conformations XIV and XV of these compounds. Such conformations



(XIII)



(XIV)



(XV)

for piperidines have been postulated by Barton and Cookson (19), who consider the space-filling requirements of the orbital containing the unshared electron pair to be greater than those of the hydrogen atom; and recently these conformations have been supported by measurements of molar Kerr constants (20). On these grounds one would expect the stable conformations of the N-methyl compounds (III and VIII, R = Me) to contain the structural unit XIII, since the space-filling characteristics of the methyl group will certainly be greater than those of the pair of unshared electrons. N-Methyl-4-aza-5α-cholestane (VIII, R = Me) and N-methyl-4-aza-5β-cholestane (III, R = Me) were prepared by heating the secondary amines (VIII and III, R = H) with formaldehyde and formic acid, and each showed an absorption band of medium intensity at 2775 cm^{-1} which disappeared when the perchlorate was formed. However, N-methyl amines generally show a peak in this region of the spectrum (21) which disappears on salt formation (22). Recently McKenna and Tulley (J. Chem. Soc. 945 (1960)) have described the preparation of the compounds III, R = Me, and VIII, R = Me, the former by a route differing from ours. On the basis of the products given by the Hofmann degradation of these compounds, they assign them the configurations assigned on other grounds in this paper.

EXPERIMENTAL

Where not otherwise specified, Merck acid-washed alumina was used for chromatography. Infrared spectra were obtained on a Perkin–Elmer Model 21 spectrophotometer; in listing the frequencies of absorption bands, the intensities are designated w (weak), m (medium), s (strong), sh (shoulder). Microanalyses were carried out by W. Manser (Zurich) and A. Bernhardt (Mülheim, Germany).

A-Nor-5-β-cholestan-3-one

3,4-Seco-5β-cholestane-3,4-dioic acid (2.60 g) (23), covered with 12 ml of acetic anhydride, was gradually heated, over a period of 30 minutes, to remove most of the

acetic anhydride by distillation at atmospheric pressure. The temperature was then raised and the fraction distilling at 260°–274° C/4 mm was collected as a pale yellow oil. An ether solution of the oil was washed with alkali and then water, dried (sodium sulphate), and concentrated. The residue dissolved partially in warm methanol, from which crystals (0.57 g), m.p. 72°–73° C, separated on cooling. The residue from the mother liquors was taken up in pentane and chromatographed on alumina. Elution with pentane containing increasing amounts of benzene gave, on concentration, a further crop (1.00 g), m.p. 72°–73° C. Recrystallization of the combined crops from methanol afforded 1.35 g (60% yield) of A-nor-5β-cholestan-3-one, m.p. 73°–74° C, $[\alpha]_D^{24}$ +122° (*c*, 1.995 in chloroform) (lit. m.p. 73°–74° C (8), $[\alpha]_D$ +105° (chloroform) (24)).

A-Nor-5β-cholestan-3-one Oxime (I)

A mixture of A-nor-5β-cholestan-3-one (0.60 g), hydroxylamine hydrochloride (0.5 g), and anhydrous sodium acetate (0.6 g) in 75% aqueous ethanol (50 ml) was refluxed for 4 hours. The precipitate obtained on dilution with water was crystallized from methanol to give rosettes of the *oxime* (0.55 g), m.p. 123°–124° C, $[\alpha]_D^{24}$ +81.7° (*c*, 0.416 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3630 m, 3290 m, 3170 sh, 2940 s, 2860 sh, 1617 w, 1465 s, 1440 m, 1425 sh, 1377 m, 1364 w, 1342 w, 1319 w, 1304 w, 949 m, 919 m, 888 m. Calc. for $C_{22}H_{45}ON$: C, 80.55; H, 11.71; N, 3.62%. Found: C, 80.57; H, 11.80; N, 3.62%.

4-Aza-5β-cholestan-3-one (II)

A solution of A-nor-5β-cholestan-3-one oxime (0.50 g) and *p*-acetylaminobenzenesulphonyl chloride (0.25 g) in absolute pyridine (13 ml) was allowed to stand for 2 days (cf. 25). The solution was then diluted with water and acidified with dilute sulphuric acid. The pale brown solid (0.49 g) thus obtained contained lactam (strong infrared band at 1660 cm^{-1}) and oxime (bands at 949, 919, and 888 cm^{-1}), and melted at 119–135° C. After several recrystallizations from methanol, the material (258 mg) was chromatographed on alumina (6 g, activity I). Elution with ether–acetone and acetone removed starting material (173 mg). Further elution with ethanol:acetone (1:2, 100 ml) afforded impure 4-aza-5β-cholestan-3-one (59 mg), purified by repeated recrystallization from methanol to give the lactam (19 mg), m.p. 151–155° with decomposition (bath preheated to 140° C), $[\alpha]_D^{24}$ +68.2° (*c*, 0.909 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3410 w, 3205 w, 3080 w, 2940 s, 2870 sh, 1660 s, 1460 m, 1443 m, 1408 w, 1377 m, 1355 w, 1303 m, 920 w. Calc. for $C_{26}H_{45}ON$: C, 80.55; H, 11.71; N, 3.62%. Found: C, 80.40; H, 11.57; N, 3.58%.

N-Hydroxy-4-aza-5β-cholestan-3-one (VI, R = H)

The oxime of 5-oxo-3,5-A-norcholestan-3-oic acid (2.20 g), prepared according to Windaus (2), was dissolved in acetic acid (125 ml) and hydrogenated in a Parr apparatus for 4.5 hours in the presence of platinum oxide (0.50 g). After removal of the platinum, the solution was concentrated at reduced pressure to give a brownish residue. Recrystallization of this material from methanol gave a first crop (1.30 g) of white crystals of N-hydroxy-4-aza-5β-cholestan-3-one; a second crop (0.35 g) was obtained from the mother liquors. After two more crystallizations, the melting point was 122°–124° C, $[\alpha]_D^{24}$ +31.8° (*c*, 1.949 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3240 w, 2940 s, 2870 sh, 1623 s, 1465 m, 1443 sh, 1430 sh, 1388 m, 1360 sh, 1328 w, 1314 sh, 1299 m, 1165 w, 1140 w, 1095 w. With chloroform as solvent, the carbonyl absorption occurred at 1607 cm^{-1}. In both carbon tetrachloride and chloroform the carbonyl frequency was not changed by dilution of the solution. The compound gave a reddish-purple color in alcoholic ferric chloride

solution. When the compound was dried at 65°/0.1 mm, it analyzed as a monohydrate. Calc. for $C_{26}H_{47}O_2N$: C, 74.07; H, 11.23; N, 3.32%. Found: C, 74.25, 74.14; H, 11.13, 11.14; N, 3.33%. When dried at 110°/0.1 mm for 4 hours the water of crystallization was lost. Calc. for $C_{26}H_{45}ON$: C, 77.40; H, 11.24; N, 3.47%. Found: C, 77.33; H, 11.28; N, 3.55, 3.48%.

N-Benzoyloxy-4-aza-5β-cholestan-3-one (VI, R = PhCO)

A solution of N-hydroxy-4-aza-5β-cholestan-3-one (200 mg) and benzoyl chloride (1.5 ml) in absolute pyridine (3.0 ml) was refluxed for 1 hour. After the reaction mixture was cooled, it was diluted with ether and washed with aqueous sodium bicarbonate and then with water. After it had been dried (magnesium sulphate) and concentrated the ether produced a yellow oil which was chromatographed on alumina (5 g, Woelm, activity II–III). Elution with benzene–ether (1:1, 50 ml; 1:3, 50 ml) gave 40 mg of crystalline material, which, after one recrystallization from ligroin–benzene, had a melting point of 209°–212° C, $[\alpha]_D^{24}$ −6.4° (*c*, 0.621 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3420 w, 3220 w, 3080 w, 2940 s, 2870 sh, 1720 s, 1677 s, 1565 m, 1450 m, 1443 sh, 1425 sh, 1377 m, 1357 w, 1330 w, 1312 m, 1265 s, 1174 m, 1110 s, 1093 sh, 1065 m, 1025 m, 902 w. Calc. for $C_{33}H_{49}O_3N$: C, 78.06; H, 9.73; N, 2.76%. Found: C, 77.55; H, 9.66; N, 3.10%.

N-Methoxy-4-aza-5β-cholestan-3-one (VI, R = Me)

A mixture of N-hydroxy-4-aza-5β-cholestan-3-one (475 mg), methyl iodide (0.35 ml), and 7% aqueous potassium hydroxide (2.0 ml) in ethanol (10 ml) was refluxed for 4.5 hours. The alcohol was removed under reduced pressure, the mixture was diluted with aqueous 10% potassium hydroxide, and extracted several times with ether. The ether extract, worked up in the usual way, gave a white solid (278 mg). Part of the substance (95 mg) was chromatographed over alumina (3.0 g, Woelm, activity II). Elution with benzene–hexane (4:1, 20 ml) and benzene (80 ml) gave crystalline material (70 mg), m.p. 106°–108° C. One recrystallization from aqueous methanol gave plates, m.p. 114°–115° C, $\nu_{max}^{CCl_4}$ in cm^{-1}: 2945 s, 2870 sh, 1688 sh, 1658 s, 1470 m, 1445 m, 1385 m, 1333 w, 1295 m, 1245 w, 1220 w, 1097 w, 1051 m, 955 w, 955 w, 909 w, 892 w, 850 w. Calc. for $C_{27}H_{48}O_2N$: C, 77.46; H, 11.55; N, 3.35%. Found: C, 77.69; H, 11.43; N, 3.40%.

4-Aza-5β-cholestane (III, R = H)

(a) By the Reduction of 4-Aza-5β-cholestan-3-one

A mixture of lithium aluminum hydride (0.30 g) and 4-aza-5β-cholestan-3-one (309 mg) in dioxane (35 ml) was refluxed for 40 hours. The mixture was then treated with water, and extracted four times with chloroform. The extracts were dried over magnesium sulphate and evaporated to give 4-aza-5β-cholestane as an uncrystallizable oil, $[\alpha]_D^{33}$ +22.1° (*c*, 1.808 in chloroform), ν_{max}^{film} in cm^{-1}: 3305 w, 2940 s, 2870 sh, 2800 sh, 2730 sh, 2620 sh, 1467 s, 1445 s, 1382 s, 1337 w, 1312 m, 1254 w, 1234 w, 1214 w, 1169 w, 1128 m, 1108 sh, 1067 w, 1030 w, 988 w, 968 w, 931 w, 884 w, 860 w, 837 w, 806 w, 777 w, 724 w. The *perchlorate* was precipitated by adding a few drops of 70% perchloric acid to 4-aza-5β-cholestane (50 mg) in methanol (5 ml). Crystallized from methanol, it had a melting point of 203–206° C, $[\alpha]_D^{33}$ +21.4° (*c*, 2.049 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3190 sh, 3100 m, 2950 s, 2870 sh, 1590 w, 1467 m, 1455 sh, 1430 w, 1407 w, 1381 m, 1366 sh, 1336 w, 1309 w, 1287 w, 1125 s, 1090 s, 918 w, 903 sh, 856 w. Calc. for $C_{26}H_{48}O_4NCl$: C, 65.85; H, 10.20; N, 2.95%. Found: C, 65.70; H, 10.29; N, 2.99%.

(b) By the Reduction of N-Hydroxy-4-aza-5β-cholestan-3-one

A mixture of granulated tin (3.0 g) and N-hydroxy-4-aza-5β-cholestan-3-one (307 mg)

in ethanol (9 ml) and concentrated hydrochloric acid (3 ml) was refluxed on the steam bath for 7 hours. Every hour a further 1 ml of hydrochloric acid was added, as well as enough ethanol to keep the organic material in solution. The solution was decanted from tin and diluted with water, giving 4-aza-5β-cholestane hydrochloride (295 mg), obtained after two recrystallizations from acetone as short needles which darkened at 237° and melted at 256°–265° C (bath preheated to 230° C). Calc. for $C_{26}H_{48}NCl$: C, 76.14; H, 11.80%. Found: C, 75.92; H, 11.76%.

The free base, obtained in the usual way as an uncrystallizable oil, $[\alpha]_D^{33}$ +24.2 (*c*, 0.910 in chloroform), was converted by treatment with 70% perchloric acid to a crystalline perchlorate, which, after one recrystallization from methanol, had a melting point of 201°–205° C, $[\alpha]_D^{33}$ +19.7° (*c*, 1.521 in chloroform). This compound was shown by mixed melting point and by its infrared spectrum to be identical with the perchlorate obtained by reduction of 4-aza-5β-cholestan-3-one.

N-*Methyl-4-aza-5β-cholestane* (III, R = Me)

A mixture of 4-aza-5β-cholestane (97 mg), 40% formaldehyde solution (1.0 ml), and 99% formic acid (1.0 ml) was heated on the steam bath for 3 hours. After the mixture was cooled, it was diluted with water and made basic with ammonium hydroxide. Extraction with ether and working up in the usual way gave a colorless uncrystallizable syrup (90 mg), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 s, 2870 sh, 2775 m, 2715 sh, 2620 w, 1465 s, 1445 s, 1376 s, 1263 w, 1236 w, 1215 w, 1200 w, 1187 w, 1173 w, 1143 w, 1112 w, 1090 w, 1073 w, 1025 sh, 1014 m, 978 w, 960 w, 930 w, 877 w, 833 w. The *perchlorate* had a melting point of 246°–251° C; in its infrared spectrum the absorption band at 2775 cm^{-1} had disappeared and bands at 1100 and 1052 cm^{-1} (perchlorate anion) had appeared. Calc. for $C_{27}H_{50}O_4NCl$: C, 66.40; H, 10.32; N, 2.87%. Found: C, 66.60; H, 10.39; N, 3.04%.

4-Aza-5α-cholestan-3-one (VII)

Ammonium formate (2.5 g) was distilled from a 25-ml conical flask until a temperature of 165° C was reached in the flask. 5-Oxo-3,5-seco-A-norcholestan-3-oic acid (1.70 g) suspended in nitrobenzene (10 ml) was introduced into the flask and the temperature of the resulting mixture was kept at 180°–200° C for 24 hours (cf. 26). After the reaction mixture was cooled, it was washed with water. The water-insoluble liquid was refluxed for 2 hours in 6 *N* hydrochloric acid (8 ml) and ethanol (7 ml). After the removal of nitrobenzene and ethanol by steam distillation, the aqueous solution was extracted with ether. The ether solution was washed with dilute sodium carbonate and with water, dried over sodium sulphate, and evaporated to dryness. The residue, after one recrystallization from methanol, weighed 1.27 g, m.p. 240°–251° C. Several recrystallizations from methanol raised the melting point to 250°–252° C with decomposition (bath preheated to 240° C), $[\alpha]_D^{22}$ +45° (*c*, 1.585 in chloroform), (lit. m.p. 252°–253° C, $[\alpha]_D$ +44 (in pyridine) (1)). The melting point and infrared spectrum of this compound were identical with those of authentic 4-aza-5α-cholestan-3-one obtained by the reduction of the oxime of 5-oxo-3,5-seco-A-norcholestan-3-oic acid (V) with sodium and alcohol (1); $\nu_{max}^{CCl_4}$ in cm^{-1}: 3422 w, 3210 w, 3070 w, 2940 s, 2870 sh, 1662 s, 1467 s, 1444 sh, 1404 m, 1383 m, 1358 m, 1306 m, 1223 w, 1162 w, 1132 w. Calc. for $C_{26}H_{45}ON$: N, 3.62%. Found: N, 3.70%.

N-*Hydroxy-4-aza-5α-cholestan-3-one* (IX)

The oxime of 5-oxo-3,5-seco-A-norcholestan-3-oic acid (0.64 g) and zinc dust (1.09 g) in acetic acid (20 ml) were heated to 90–95° C for 1 hour. The zinc was then removed

by filtration, and the filtrate diluted with water to yield N-hydroxy-4-aza-5α-cholestan-3-one as a fluffy white solid (0.59 g, m.p. 140°–153° C). Repeated recrystallizations from ethyl acetate and then from methanol raised the melting point to 185°–187° C with decomposition (bath preheated to 150° C), $[\alpha]_D^{28}$ −49.5° (*c*, 1.98 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3260 w, 2940 s, 2870 sh, 1677 sh, 1630 s, 1487 s, 1461 s, 1448 sh, 1428 w, 1395 s, 1380 s, 1328 w, 1315 w, 1306 m, 1267 w, 1223 s, 1187 w, 1138 w, 1072 w, 940 m, 925 w, 887 w, 860 w. Calc. for $C_{26}H_{45}O_2N$: C, 77.40; H, 11.24; N, 3.47%. Found: C, 77.23; H, 10.99; N, 3.39%. In methanolic ferric chloride the compound gave a reddish-purple color.

4-Aza-5α-cholestane (VIII, R = H)

(a) By the Reduction of 4-Aza-5α-cholestan-3-one

A mixture of 4-aza-5α-cholestan-3-one (509 mg) and lithium aluminum hydride (0.50 g) in dioxane (75 ml) was refluxed for 40 hours. Isolation of the product in the usual way gave a solid, melting at 114°–115° C, after two recrystallizations from acetone (lit. m.p. 116°–117° C (1)), $[\alpha]_D$ +36° (chloroform), ν_{max}^{KBr} in cm^{-1}: 3270 w, 2940 s, 2870 sh, 1470 s, 1445 s, 1384 s, 1368 sh, 1210 w, 1173 w, 1128 m, 1116 sh, 1077 w, 946 w, 933 m, 905 w, 889 w, 867 w, 858 w, 840 w. Calc. for $C_{26}H_{47}N$: C, 83.56; H, 12.68; N, 3.75%. Found: C, 83.44; H, 12.69; N, 3.71%.

The *perchlorate*, crystallized from methanol, had a melting point of 224°–225° C, $[\alpha]_D^{33}$ +21.1 (*c*, 2.36 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3190 sh, 3100 m, 2950 s, 2870 sh, 1590 w, 1467 m, 1450 m, 1392 sh, 1383 w, 1377 sh, 1366 sh, 1330 w, 1307 w, 1125 s, 1090 s, 918 w. Calc. for $C_{26}H_{48}O_4NCl$: C, 65.85; H, 10.20; N, 2.95%. Found: C, 65.67; H, 10.14; N, 2.72%.

(b) By the Reduction of N-Hydroxy-4-aza-5α-cholestan-3-one

A mixture of N-hydroxy-4-aza-5α-cholestan-3-one (121 mg) and lithium aluminum hydride (0.30 g) in dry ether (50 ml) was refluxed for 12 hours. Then water (0.60 ml) and 10% aqueous sodium hydroxide (0.48 ml) were added, and the mixture stirred for 10 minutes. The inorganic salts were removed by filtration, and the filtrate was evaporated to dryness. The brown syrupy residue (93 mg) had an infrared spectrum which showed no carbonyl peak. When treated with 70% perchloric acid it gave a salt, m.p. 222°–224° C, which was shown by its infrared spectrum and by a mixed melting point to be identical with the perchlorate of 4-aza-5α-cholestane obtained from 4-aza-5α-cholestan-3-one.

N-*Methyl-4-aza-5α-cholestane* (VIII, R = Me)

4-Aza-5α-cholestane (95 mg) was treated with 99% formic acid (1.0 ml) and 40% formaldehyde solution (1.0 ml) under the conditions described for preparing N-methyl-4-aza-5β-cholestane. The product (77 mg) after two crystallizations from methanol had a melting point of 63°–64° C (lit. m.p. 61°–62° C (12)), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2950 s, 2870 sh, 2775 m, 2715 sh, 2620 w, 1465 s, 1445 s, 1378 m, 1365 sh, 1328 w, 1179 m, 1160 m, 1107 m, 1068 w, 1022 m, 1000 sh, 970 sh. Calc. for $C_{27}H_{49}N$: C, 83.64; H, 12.74%. Found: C, 83.82; H, 12.97%.

The *perchlorate* melted at 231°–235° C, and had no absorption band at 2775 cm^{-1}. Calc. for $C_{27}H_{50}O_4NCl$: C, 66.40; H, 10.32; N, 2.87%. Found: C, 66.18; H, 10.24; N, 2.93%.

N-*p-Nitrobenzoyl-4-aza-5α-cholestane*

A solution of 4-aza-5α-cholestane (330 mg) and *p*-nitrobenzoyl chloride (0.25 g) in dry benzene (15 ml) was refluxed for 1.5 hours. After the solution was cooled, it was

washed with dilute sodium carbonate, dilute hydrochloric acid, and water, dried over sodium sulphate, and evaporated to dryness. Crystallization of the residue from methanol gave 180 mg of substance melting at 232°–238° C; after several recrystallizations the melting point rose to 243°–245° C (decomp.), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 s, 2865 sh, 1647 s, 1600 m, 1520 w, 1465 m, 1434 s, 1382 sh, 1363 sh, 1346 s, 1308 m, 1287 m, 1175 w, 1140 w, 1117 w, 1103 w, 1010 w, 945 w, 927 w, 856 w, 848 w. Calc. for $C_{33}H_{51}O_3N_2$: C, 75.67; H, 9.82; N, 5.32%. Found: C, 75.55; H, 9.72; N, 5.32%.

4-Azacholest-4-ene (XI)

A solution of 4-aza-5α-cholestane (0.51 g) and mercuric acetate (1.78 g) in 5% aqueous acetic acid (10 ml) was heated on the steam bath for 2 hours. After the solution was cooled, mercurous acetate was removed by filtration. The filtrate was saturated with hydrogen sulphide, centrifuged, and the supernatant liquid made strongly basic with aqueous 40% sodium hydroxide solution. The solid which precipitated was taken up in ether and the ether solution washed with water and dried (magnesium sulphate). Evaporation of the ether left a semisolid residue (0.33 g) which was dissolved in methanol (5 ml) and treated with a few drops of 70% perchloric acid. The *perchlorate* precipitated, and after one recrystallization from methanol had a melting point of 212°–213° C, ν_{max}^{KBr} in cm^{-1}: 3240 w, 3120 w, 2950 s, 2870 sh, 1680 s, 1468 s, 1455 sh, 1382 m, 1367 sh, 1335 w, 1283 w, 1213 w, 1105 s, 1080 s, 927 w. In carbon tetrachloride the C=N absorption occurred at 1677 cm^{-1}. Calc. for $C_{26}H_{46}O_4NCl$: C, 66.15; H, 9.82; N, 2.97; Cl, 7.51%. Found: C, 65.93; H, 9.68; N, 2.97; Cl, 7.34%.

The *free base* obtained from the perchlorate had a melting point of 101°–102° C after two crystallizations from acetone, $\nu_{max}^{CCl_4}$ in cm^{-1}: 2945 s, 2870 sh, 2800 sh, 2720 sh, 1649 s, 1467 s, 1445 s, 1383 m, 1377 sh, 1366 m, 1328 w, 1320 w, 1305 w, 1166 w, 1127 m, 1075 w, 1054 w, 1027 w, 954 w, 928 w, 854 w. In potassium bromide pellet also, no absorption occurred in the region 3100–3450 cm^{-1}. Calc. for $C_{26}H_{45}N$: C, 84.04; H, 12.21; N, 3.77%. Found: C, 83.78; H, 12.18; N, 3.79%.

Reduction of 4-Aza-cholest-4-ene Perchlorate

(a) With Lithium Aluminum Hydride

A slurry of 4-azacholest-4-ene perchlorate (30 mg) and lithium aluminum hydride (0.10 g) in anhydrous ether (8 ml) was refluxed with stirring for 4.5 hours, after which the excess of lithium aluminum hydride was destroyed with 30% aqueous sodium hydroxide. The aqueous solution was extracted three times with ether, and the combined ether extracts dried and concentrated to give a brownish solid (21 mg). The perchlorate of this material melted at 222°–224° C and was shown by mixed melting point and infrared spectrum to be identical with 4-aza-5α-cholestane perchlorate.

(b) By Catalytic Hydrogenation

A solution of 4-azacholest-4-ene perchlorate (92 mg) in ethanol was hydrogenated at atmospheric pressure in the presence of platinum oxide (35 mg). The calculated amount of hydrogen was taken up in 8 minutes. Removal of the catalyst by filtration, and basification and concentration of the filtrate gave a white solid, crystallizing from acetone as needles, m.p. 113°–114° C. The compound was shown by mixed melting point and infrared spectrum to be 4-aza-5α-cholestane.

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STEREOCHEMICAL STUDIES

II. CONFIGURATIONS OF SOME 4-OXACHOLESTANES[1]

J. T. EDWARD AND PETER F. MORAND[2]

ABSTRACT

The oxidation of A-nor-5β-cholestan-3-one with peracids gave 4-oxa-5β-cholestan-3-one. This was identical with one of the two lactones produced by the reduction of Windaus' keto acid; the other lactone must be 4-oxa-5α-cholestan-3-one. The two lactones were reduced by lithium aluminum hydride to diols, which were cyclized to 4-oxa-5α- and -5β-cholestane.

Turner (1) obtained by catalytic reduction of Windaus' keto acid (I) (2) two lactones, m.p. 109.5°–110° C and 116°–116.5° C, the former in larger amount. He tentatively assigned to these compounds the structures V and VI, respectively, because of their optical rotations and because hydrogenation of 4-oxacholest-5-en-3-one, which should be expected to give VI, gave the higher melting isomer.

Further evidence for these structures comes from a consideration of their molecular rotations (3) including the differences between the molecular rotations of the lactones and of the amides (II or III, R = NH_2) or hydrazides (II or III, R = $NHNH_2$), which are, as shown in Table I, in the directions predicted by the Klyne–Hudson "lactone rule"

TABLE I

Molecular rotation $[M]_D$ values

Lactones	Amides	Δ	Hydrazides	Δ
V	II, R = NH_2		II, R = $NHNH_2$	
+80°,[a] +70°[b]	+92°[a]	−12°[a]	+93°[b]	−23°[b]
VI	III, R = NH_2		III, R = $NHNH_2$	
+325°,[a] +311°[b]	+124°[a]	+201°[a]	+161°[b]	+150°[b]

[a]In 95% ethanol. [b]In chloroform.

(3). Finally, unequivocal proof comes from the Baeyer–Villiger oxidation of Windaus norketone (IV), which is known to have the cis ring juncture shown (4). With peracetic acid this affords the lactone with the lower melting point. The Baeyer–Villiger reaction proceeds with retention of configuration of migrating optically active groups (5, 6) and almost always, as in the present case, involves the migration of the more alkylated of the two carbon atoms flanking the carbonyl group (7, 8).

A notable exception to the latter generalization is afforded by the oxidation of camphor, carried out by Baeyer and Villiger themselves (9). In reinvestigating this reaction, Sauers (10) found that in acetic acid camphor is oxidized by peracetic acid to the expected lactone (XI) (82% yield) but in the presence of concentrated sulphuric acid is oxidized to campholide (XII) (30% yield), although some of the isomeric lactone (XI) also may be formed but destroyed in the strongly acidic solution. Sauers explained the formation of campholide by a mechanism involving a protonated form of the Criegee intermediate (11, 12). In the destruction of this intermediate the electronic effects which favor the migration of the more heavily alkylated carbon atom become less important than they are in the destruction of the neutral Criegee intermediate. Conformational and

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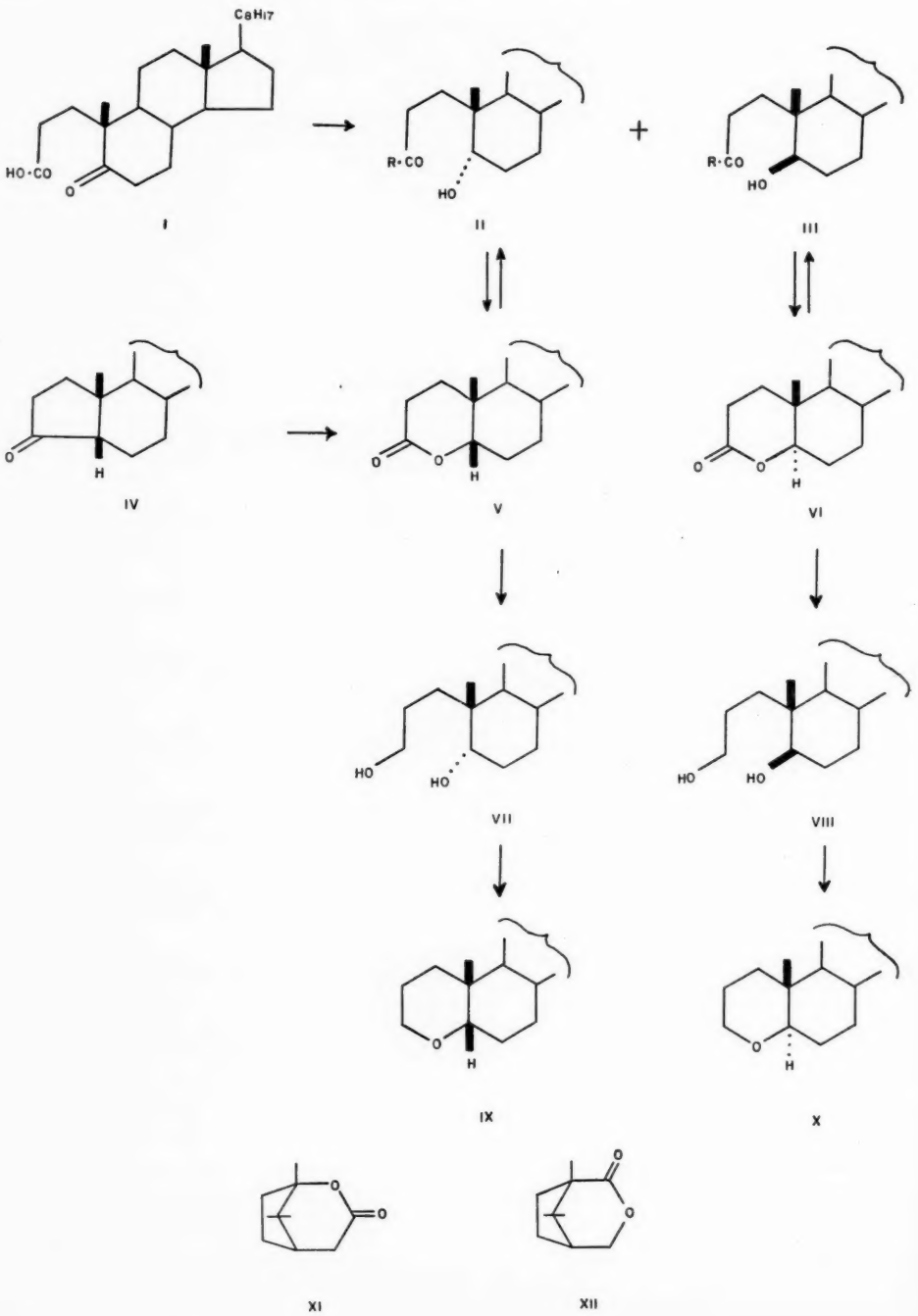
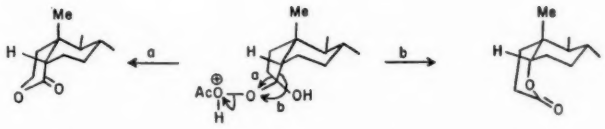


I II III IV V VI VII VIII IX X

XI XII

other effects may then become relatively more important. It had already been argued by Murray *et al.* (13) that conformational effects favor the formation of campholide since the transition state leading to it has a chair conformation, while the transition state leading to the lactone (XI) has a boat conformation.

Serious doubt is cast on this argument by the fact that A-nor-5β-cholestan-3-one (IV) even in the presence of sulphuric acid is oxidized to 4-oxa-5β-cholestan-3-one (V = XV). The line of reasoning developed by Murray *et al.* (13) would indicate that in this reaction the protonated form of the Criegee intermediate (XIV) would have a greater tendency to rearrange to give 3-oxa-5β-cholestan-4-one (XIII): indeed, the conformational effects favoring XIII over XV would appear to be greater than those favoring XII over XI.



XIII XIV XV (=V)

We have found the most convenient method of preparing the trans lactone (VI) to be the reduction of Windaus' keto acid (I) with sodium borohydride, but the yield of pure product (35%) is not great. Bolt (14) obtained this lactone by carrying out the reduction with sodium in alcohol, but his reported yield of 75% is based on the crude product; we obtained a satisfactory product, m.p. 115°–116° C, in only 34% yield. Reduction with sodium in isopropanol (15) gave a slightly higher yield (41%). Reduction with aluminum isopropoxide in isopropanol gave an oily product shown by infrared spectroscopy to contain both cis and trans lactones.

Reduction of the lactones (V and VI) with lithium aluminum hydride gave the diols (VII and VIII), from which the 4-oxacholestanes (IX and X) were obtained by treatment with benzenesulphonyl chloride and pyridine or lutidine (16). As expected, the 4-oxacholestane having the trans fusion of rings A and B had a higher melting point than the cis compound; however, the trans 4-oxa- and 4-aza-cholestanes had higher molecular rotations than the cis compounds, unlike their carbocyclic analogue (Table II). Replacing the carbon atom at position 4 by oxygen had a much greater effect on optical rotation than replacing it by nitrogen; this may be due to the greater alteration in molecular shape caused by substitution by the smaller oxygen atom, since the polarizabilities of the two atoms are not greatly different (17).

TABLE II

Physical properties of cholestanes and 4-heterocholestanes

	5α-Compounds		5β-Compounds	
4-Group	m.p. (°C)	$[M]_D$	m.p. (°C)	$[M]_D$
—CH_2—	80 (18)	+91° (3)	70 (18)	+97° (3)
—NH—	115 (19)	+134° (19)	(Oil) (19)	+82° (19)
—O—	90	+158°	52	+14°

EXPERIMENTAL

4-Oxa-5β-cholestan-3-one (V)

A solution of peracetic acid (20) in glacial acetic acid (4.0 ml containing 0.46 g of peracetic acid) (21) and *p*-toluenesulphonic acid (0.02g) were added to A-nor-5β-cholestan-3-one (0.20 g) dissolved in 2.0 ml of glacial acetic acid. The solution was left in the dark for 24 hours at 35° C. Dilution with water precipitated a product which, after one recrystallization from dilute methanol, yielded 0.17 g (90% of the theoretical) of transparent plates, m.p. 90°–99° C. Repeated recrystallizations from dilute methanol raised the melting point to 107°–107.5° C, $[\alpha]_D^{24}$ +18.6° (*c*, 6.129 in chloroform) (lit., m.p. 109.5°–110° C, $[\alpha]_D$ +18.3° (1)). The infrared spectrum of this compound was identical with that of an authentic sample of 4-oxa-5β-cholestan-3-one obtained by catalytic hydrogenation of 5-oxo-3,5-seco-A-norcholestan-3-oic acid (1), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 (s), 2870 (sh), 1738 (s), 1470 (s), 1432 (sh), 1425 (sh), 1384 (m), 1367 (m), 1345 (m), 1325 (sh), 1235 (w), 1198 (m), 1173 (w), 1156 (m), 1113 (w), 1095 (w), 1075 (w), 1040 (s), 1016 (sh) (s = strong; m = medium; w = weak band; sh = shoulder). Calc. for $C_{26}H_{44}O_2$: C, 80.38; H, 11.41%. Found: C, 80.24; H, 11.24%.

In another experiment, a solution of the norketone (205 mg) in glacial acetic acid (2 ml) and concentrated sulphuric acid (0.75 ml) was treated with a solution of peracetic acid in glacial acetic acid (0.75 ml 40%). The solution was left in the dark at 27° for 5 days, and then worked up as before. The product crystallized once from aqueous ethanol gave plates (55 mg), m.p. 85°–95°; after several recrystallizations the melting point was raised to 107°–108° C, and the product was shown by mixed melting point and infrared spectrum to be 4-oxa-5β-cholestan-3-one. No other crystalline product could be isolated from the mother liquors.

4-Oxa-5α-cholestan-3-one (VI)

5-Oxo-3,5-seco-A-norcholestan-3-oic acid (2.5 g) (1) was dissolved in 30 ml of 0.8 *N* sodium hydroxide and 7.5 ml of ethanol. Sodium borohydride (0.29 g) was added over a period of 10 minutes and the mixture was heated under reflux for 1.5 hours. The excess sodium borohydride was destroyed with acetic acid and the acidified reaction mixture was refluxed for 10 minutes longer to bring about complete lactonization of the hydroxy acid. After the precipitated product was filtered, several recrystallizations from petroleum ether (b.p. 66°–75° C) yielded 0.90 g of 4-oxa-5α-cholestan-3-one, m.p. 114°–115° C, $[\alpha]_D^{24}$ +80.0° (*c*, 1.526 in chloroform) (lit., m.p. 116°–116.5° C, $[\alpha]_D$ +80.5° (1)), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 (s), 2870 (sh), 1740 (s), 1467 (m), 1448 (sh), 1392 (sh), 1383 (w), 1372 (sh), 1342 (w), 1186 (m), 1130 (w), 1077 (m), 1047 (s). The residue from the mother liquors (0.96 g) was dissolved in benzene and chromatographed on alumina (24 g). Elution with benzene (90 ml) and benzene–acetone (19:1, 175 ml) gave 0.46 g of material with melting range 78°–106° C. Several recrystallizations from dilute methanol yielded 4-oxa-5β-cholestan-3-one (0.11 g), m.p. 107°–107.5° C, $[\alpha]_D^{24}$ +18.1°.

5β-Hydroxy-3,5-seco-A-norcholestan-3-amide (III, R = NH_2)

A solution of 4-oxa-5α-cholestan-3-one (VI) (0.13 g) in methanol saturated with ammonia deposited plates (0.12 g) of 5β-hydroxy-3,5-seco-A-norcholestan-3-amide, which after recrystallizing from methanol had a melting point of 149°–155° C (in a bath preheated to 135° C), $[\alpha]_D^{23}$ +30.5° (*c*, 0.918 in 95% ethanol). Calc. for $C_{26}H_{47}O_2N$: C, 76.98; H, 11.68; N, 3.45%. Found: C, 76.41; H, 11.33; N, 3.10%. The compound decomposed at its melting point (which was unsharp and dependent on the rate of heating) to regenerate the lactone (VI), identified by melting point, mixed melting point, and infrared spectrum.

5α-Hydroxy-3,5-seco-A-norcholestan-3-amide (II, R = NH_2)

Prepared in the same manner from the cis lactone (V), the amide melted at 153°–159° C (bath preheated to 135° C), $[\alpha]_D^{23}$ +22.6° (*c*, 1.06 in 95% ethanol). Calc. for $C_{26}H_{47}O_2N$: N, 3.45%. Found: N, 3.21%.

5β-Hydroxy-3,5-seco-A-norcholestan-3-hydrazide (III, R = $NHNH_2$)

A solution of 4-oxa-5α-cholestan-3-one (0.25 g) and 95% hydrazine (0.25 ml) in methanol (5 ml) after 15 hours deposited needles of the hydrazide, m.p. 168°–170° C (bath preheated to 160° C), $[\alpha]_D^{24}$ +38.4° (*c*, 2.70 in chloroform). Calc. for $C_{26}H_{48}O_2N_2$: C, 74.25; H, 11.50; N, 6.66%. Found: C, 74.27; H, 11.56; N, 6.75%.

5α-Hydroxy-3,5-seco-A-norcholestan-3-hydrazide (III, R = $NHNH_2$)

Prepared in the same way from the cis lactone (V), the hydrazide melted at 178°–181° C (bath preheated to 160° C), $[\alpha]_D^{24}$ +22.0° (*c*, 0.910 in chloroform). Calc. for $C_{26}H_{48}O_2N_2$: C, 74.25; H, 11.50; N, 6.66%. Found: C, 74.02; H, 11.56; N, 6.63%.

3,5-Seco-A-norcholestan-3,5α-diol (VII)

A solution of 4-oxa-5β-cholestan-3-one (454 mg) in ether (10 ml) was added to a solution of lithium aluminum hydride (376 mg) in ether (25 ml) in 10 minutes with stirring. The mixture was refluxed for 2.5 hours, and then water (1.0 ml) and 10% aqueous sodium hydroxide (0.8 ml) were added dropwise with continued stirring. The precipitated inorganic salts were removed by filtration. Concentration of the ether filtrate and washings gave white flakes (396 mg), m.p. 121°–127° C, raised by several recrystallizations from aqueous methanol to a melting point of 130°–131° C, $[\alpha]_D^{28}$ +45.7 (*c*, 0.656 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3655 (w), 3360 (m), 2940 (s), 2870 (sh), 1468 (m), 1445 (m), 1380 (m), 1368 (sh), 1112 (w), 1053 (m), 1032 (sh), 995 (w), 960 (w), 929 (w), 917 (w). Calc. for $C_{26}H_{48}O_2$: C, 79.50; H, 12.32%. Found: C, 79.20; H, 12.14%.

4-Oxa-5β-cholestane (IX)

Benzene sulphonyl chloride (0.09 ml) was added over a period of 5 minutes to a refluxing solution of 3,5-seco-A-norcholestan-3,5α-diol (232 mg) in 2,6-lutidine (0.21 ml). The mixture was cooled, diluted with water (15 ml), and extracted with ether. Removal of the ether left a brownish semisolid residue (210 mg), which in petroleum ether (b.p. 35–60°, 100 ml) was passed through an alumina column (5.0 g, Woelm, activity III) to give waxy crystals (107 mg), m.p. 39–46°. Rechromatography on alumina (3.5 g, Woelm, activity I), followed by crystallization from methanol gave rosettes of white needles, m.p. 51°–52° C, $[\alpha]_D^{27}$ +3.8° (*c*, 2.10 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 (s), 2860 (sh), 2720 (w), 1466 (s), 1443 (s), 1377 (s), 1365 (sh), 1356 (sh), 1340 (w), 1328 (w), 1250 (w), 1204 (w), 1193 (w), 1165 (w), 1155 (w), 1113 (sh), 1098 (s), 1086 (s), 1067 (sh), 1026 (m), 982 (w), 958 (w), 937 (sh), 932 (w), 905 (w), 880 (m), 857 (w). Calc. for $C_{26}H_{46}O$: C, 83.36; H, 12.37%. Found: C, 83.23; H, 12.25%.

3,5-Seco-A-norcholestan-3,5β-diol (VIII)

A mixture of 4-oxa-5α-cholestan-3-one (0.50 g) and lithium aluminum hydride (0.06 g) in tetrahydrofuran (50 ml) was stirred for 4 hours at room temperature, and then treated with water (50 ml) and 1 *N* sulphuric acid (5.0 ml). Extraction with ether followed by removal of the ether gave a gelatinous material (0.39 g), m.p. 128°–131° C. One recrystallization from methanol raised the melting point to 134°–135° C, $[\alpha]_D^{18}$ +25.3° (*c*, 1.61 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 3655 (w), 3395 (m), 2940 (s), 2870 (sh), 1467 (m), 1446 (m), 1383 (m), 1366 (sh), 1338 (w), 1082 (w), 1051 (m), 1025 (m), 1000 (sh), 950 (w), 925 (w), 905 (w). Calc. for $C_{26}H_{48}O_2$: C, 79.50; H, 12.32%. Found: C, 79.38; H, 12.62%.

4-Oxa-5α-cholestane (X)

Benzene sulphonyl chloride (0.19 ml) was added over 5 minutes to a refluxing solution of 3,5-seco-A-norcholestan-3,5β-diol (0.51 g) in absolute pyridine (0.45 ml). The solution was refluxed for another 5 minutes, cooled, diluted with water (50 ml), and extracted with ether. The ether extracts were washed with dilute sulphuric acid and then with water, dried (Na_2SO_4), and concentrated to give a crude brownish product (0.33 g). After several recrystallizations from methanol this had a melting point of 89°–90° C, $[\alpha]_D^{23}$ +42.3° (*c*, 1.33 in chloroform), $\nu_{max}^{CCl_4}$ in cm^{-1}: 2940 (s), 2860 (sh), 2720 (w), 1465 (s), 1445 (s), 1383 (s), 1377 (sh), 1367 (sh), 1267 (w), 1193 (w), 1169 (w), 1154 (w), 1122 (m), 1101 (s), 1086 (s), 1040 (m), 1027 (w), 978 (w), 954 (w), 928 (m), 909 (w), 876 (w), 854 (w). Calc. for $C_{26}H_{48}O$: C, 83.36; H, 12.37%. Found: C, 82.96; H, 12.29%.

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CO-ORDINATION COMPLEXES OF TITANIUM (IV) HALIDES

1. PREPARATION AND INFRARED SPECTRA OF THE COMPLEXES OF TITANIUM (IV) HALIDES WITH N,N-DIMETHYL AND N,N-DIPHENYL FORMAMIDE[1]

JACQUES ARCHAMBAULT[2] AND ROLAND RIVEST

ABSTRACT

Co-ordination complexes of titanium (IV) halides with N,N-dimethyl formamide and N,N-diphenyl formamide have been prepared and their infrared spectra have been measured. In all these complexes, there is a metal-to-oxygen bond as evidenced in the infrared spectra by a negative shift of the carbonyl stretching frequency and a positive shift of the CN stretching frequency. The relative values of these shifts and the ease of decomposition observed for these complexes indicate that their stability decreases with decreasing the electronegativity of the halide bonded to titanium (IV).

In the complexes of N,N-diphenyl formamide, the effect of the conjugated phenyl groups has been observed. In these complexes the decrease of double-bond character of the carbonyl group and the increase of double-bond character of the CN group, resulting from the metal-to-oxygen co-ordination bond, are quite larger than in the complexes of N,N-dimethyl formamide.

INTRODUCTION

When a Lewis acid, as an example MX_4, is a covalent molecule, its behavior is analogous to that of a metallic ion M^{+4} of which the co-ordination sphere is partially filled. The acidity of the molecule depends upon the tendency of the ion to complete its co-ordination sphere. The nature of the negative ion X^- affects the acidity of the central metallic ion. If X^- increases in size and decreases in electronegativity, it will be more polarized by the central ion and will decrease the affinity of this ion for an additional electron pair.

In the case of the halides of titanium (IV), the equation of Hannay and Smith (1) and the electronegativity scale of Pauling (2) indicate a bonding predominantly ionic for titanium tetrafluoride and predominantly covalent for the other three halides. The acidity of these covalent halides decreases from the chloride to the iodide, the partial ionic character of the latter being less than 20%.

In a previous paper (3) the formation of complexes of titanium tetrachloride with formamide and N,N-dimethyl formamide was explained by a co-ordination through the oxygen of the amide. In order to see how the difference in covalent character between the tetrahalides of titanium would influence the co-ordination of amides to titanium (IV), we have extended our investigation to titanium tetrabromide and titanium tetraiodide.

If a donor molecule contains some electrons related to the donor atom by a conjugated system, some type of interactions between these electrons and the donor atom should be noticeable when the latter is used in co-ordination. We have then prepared the complexes of N,N-diphenyl formamide and we have observed that the electrons of the phenyl groups have a strong influence on the co-ordination bond.

EXPERIMENTAL

Materials and Manipulations

Reagent grade titanium tetrachloride was purified by the method of Gilchrist (4), redistilled, and the medium fraction was used for our experiments. Titanium tetrabromide, supplied by E. D. Mackay, New York, has been distilled under reduced pressure.

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The medium fraction, distilling at 100° C under 10 mm Hg, had the following composition: Ti, 13.04%; Br, 85.80%. The calculated values for pure $TiBr_4$ are: Ti, 13.03%; Br, 86.97%.

Titanium tetraiodide of E. D. Mackay Co., New York, was contaminated by free iodine. The material was first heated at 100° C under vacuum to sublimate the iodine. It was then distilled at 170–175° C under 2 mm Hg. The medium fraction had the following composition: Ti, 8.71%; I, 90.1%. The calculated value for pure TiI_4 are Ti, 8.62%; I, 91.38%.

Reagent grade of N,N-dimethyl formamide has been redistilled and the medium fraction used for our experiments was kept over dry sodium sulphate.

Reagent grade of N,N-diphenyl formamide was dried under vacuum and kept in a desiccator containing phosphorus pentoxide.

The solvents used for our experiments were anhydrous. Their water content, checked by the Karl Fischer method, was only a few parts per million.

The water-sensitive complexes were prepared and handled in a dry box continuously flushed with dry nitrogen. The dry box was provided with electricity and vacuum so that all the operations could be performed in an atmosphere of very low humidity content. Under these experimental conditions, there is always a certain amount of hydrolysis of the halide, as the results of elementary analyses indicate.

The methods of analysis for chlorine, titanium, carbon, and hydrogen were described previously (3). Bromine and iodine determinations were made by the volumetric method of Volhard. Nitrogen was determined by the Kjeldahl method or by the micro-Dumas method as specified by a "k" or a "d" in each case.

The setup used for the preparation of complexes from solutions of the reactants was identical with the one used in previous experiments (3). For the preparation of the iodide complex, solid TiI_4 was added to the amide solution.

The spectroscopic measurements were carried out on a Perkin–Elmer infrared spectrophotometer, model 21, using a sodium chloride prism. The spectra of the complexes were determined in solutions of suitable solvents.

The System $TiCl_4$, $2HCON(C_6H_5)_2$

Twenty-five milliliters of an 8.5% solution of titanium tetrachloride in carbon tetrachloride was added dropwise to 45 ml of a 13% solution of N,N-diphenyl formamide in the same solvent. There was no precipitation at the beginning, but precipitation occurred when 80% of the titanium tetrachloride was added, leading to an orange-yellow powder, melting at 119–121° C. The powder was washed many times with pure carbon tetrachloride on a fritted glass filter, and then dried *in vacuo*. This powder is very slightly soluble in carbon tetrachloride, but is soluble in aromatic hydrocarbons and in halogen derivatives of aliphatic hydrocarbons. The analysis of this adduct led to the following results.

	% Ti	% Cl	% N	% C	% H	Ti/Cl
Calc. for $TiCl_4.2HCON—(C_6H_5)_2$	8.20	24.28	4.8	53.45	3.80	1/4.00
Found	8.05	22.39	4.7d	54.74	3.73	1/3.76

The System $TiBr_4$—$HCON(C_6H_5)_2$

Fifteen milliliters of a 13% solution of diphenyl formamide in carbon tetrachloride was added dropwise to 20 ml of a 10% solution of titanium tetrabromide in the same

solvent. The adduct obtained was a red oil. By evaporation of the solvent the compound became a solid of resinous appearance. It was washed many times with normal hexane and dried *in vacuo*. The final compound was a red powder which decomposed at 60–65° C. The elementary analysis gave the following results.

	% Ti	% Br	% N	% C	% H	Ti/Br
Calc. for $TiBr_4.2HCON(C_6H_5)_2$	6.29	41.95	3.67	40.97	2.91	1/4.00
Found	6.91	40.40	3.3d	36.21	3.02	1/3.50

The System $TiBr_4.2HCON(CH_3)_2$

Thirty milliliters of a 20% solution of $TiBr_4$ in carbon tetrachloride was added dropwise to 30 ml of a 10% solution of the amide. A red oil was obtained at the beginning of the reaction and after agitation for many hours a red solid precipitated. Its melting point was 178–181° C. The elementary analyses led to the following results.

	% Ti	% Br	% N	% C	% H	Ti/Br
Calc. for $TiBr_4.2HCON(CH_3)_2$	9.32	62.22	5.45	14.03	2.75	1/4.00
Found	9.32	60.3	5.04k 5.73d	14.03	3.76	1/3.87

The System TiI_4—HCON—$(CH_3)_2$

Powdered titanium tetraiodide was added by small quantities to a solution of the amide in carbon tetrachloride. A deep brown solution was first observed and precipitation occurred. The brown powder so obtained was washed with carbon tetrachloride to remove the excess of amide. The powder was extremely unstable even in the absence of moisture. It melted at room temperature, becoming like a paste after a few days.

Two series of experiments were performed between the iodide and the solutions of the amide where the ratio of titanium tetraiodide to the amide was 1 to 3 and 1 to 4,2 respectively. In a third experiment, the iodide was added to a 10-fold excess of the pure amide, without solvent. In each experiment, the appearance of the compound obtained was the same but the analytical results were quite different.

	% Ti	% I^-	% T_2	Ti/I total
Expt. 1, found	5.76	56.8±0.5	1.5	1/3.9
Expt. 2, found	5.28	59.9±0.5	2.1	1/4.2
Calc. for $TiI_4.4HCON(CH_3)_2$	5.65	59.7	0	1/4.00
Expt. 3, found	4.21	42.	2.5	1/4
Calc. for $TiI_4.8HCON(CH_3)_2$	4.2	44.5	0	1/4

These analytical data are not necessarily indicating a higher co-ordination number for Ti(IV) but might be the result of a solvation effect of the complex $TiI_4.2HCON(CH_3)_2$ with the pure amide.

In a previous paper (5), we have demonstrated the existence of a compound of the following composition $TiCl_3OC_4H_9.6C_4H_9OH$. Fowles (6) also reported the formation of complexes of titanium halides with ammonia where the molecular formula was $TiX_4.8NH_3$. As explained by this author, it is very unlikely that the co-ordination number of titanium (IV) is higher than six but it is more probable that the extra molecules of

ligands are held in a secondary sphere of influence rather like the ammoniates formed by cobaltammine-type ion (7). The presence of free iodine in these compounds could be related to the formation of trivalent titanium. Fowles (8) has proposed that mechanism to explain the formation of iodine in the thermal decomposition of $TiI_4 . 6NH_3$.

Infrared Spectra and Structure Discussion

The Complexes of N,N-*Dimethyl Formamide*

The absorption maxima observed for the complexes of the titanium halides with N,N-dimethyl formamide are listed in Table I and illustrated in Fig. 1. The carbonyl

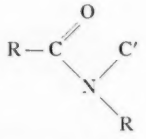
TABLE I

Frequencies of N,N-dimethyl formamide complexes

$HCON(CH_3)_2$		$TiCl.2HCON(CH_3)_2$	$TiBr_4 . 2HCON(CH_3)_2$		$TiI_4 . nHCON(CH_3)_2$	Assignment
Soln. CH_2Cl_2	Pure liquid	Soln. CH_2Cl_2	Soln. CH_2Cl_2	Nujol mulls	Soln. CH_2Cl_2	
1730 sh	1725 sh	1657 sh	1655 sh	1662 sh	1665 sh 1680 m	—
1681 vs	1675 vs	1643 vs	1647 vs	1642 vs	1653 vs	ν(C=O)
1442 w 1413		1428 w	1425 m	1412 m	1430 m	
1393 s	1393 s	1362 m	1362 m	1340 s	1367 s	
1093 m	1094 m	1118 w		1127 w	1123 w 1096	Possible ν(CN)
1064	1065 w			1053 w	1060 w	

ABBREVIATIONS: vs, very strong; m, medium; w, weak; sh, shoulder.

stretching frequency is shown to be lowered by 37, 34, and 27 cm^{-1} in the spectra of the chloride, of the bromide, and of the iodide complex respectively. Another band tentatively assigned to the C—N stretching vibration is shifted to higher frequencies by about the same value. As it was explained in a previous paper (3), no data are available concerning the assignment of the infrared absorption bands of N,N-dimethyl formamide. According to the following skeleton,



some authors consider the C—N vibrations in disubstitute amides as being confused with the C=O stretching band at about 1650 cm^{-1}. Mizushima and others (9) assigned the absorption bands at 1160 and 1040 cm^{-1} of N-methyl acetamide to the C—N and C′—N vibrations respectively. More recently, De Graff and Sutherland (10) proposed that a band at 956 cm^{-1} of N-methyl acetamide be assigned to C′—N stretching vibration. It seemed reasonable to us that a positive frequency shift of the band located at 1093 cm^{-1} when the amide co-ordinates to Ti(IV) was an argument in favor of its assignment to the C—N stretching vibration. The frequency shifts of the carbonyl band and of the C—N band are an indication of co-ordination through the oxygen atom. The shifts observed seem to decrease from the chloride complex to the iodide complex. If the magnitude of the shift can be related to the stability of the co-ordinated bond, the observed trend is in good agreement with the temperature of decomposition, which are 220–225° C for the chloride complex, 178–181° C for the bromide complex, and room

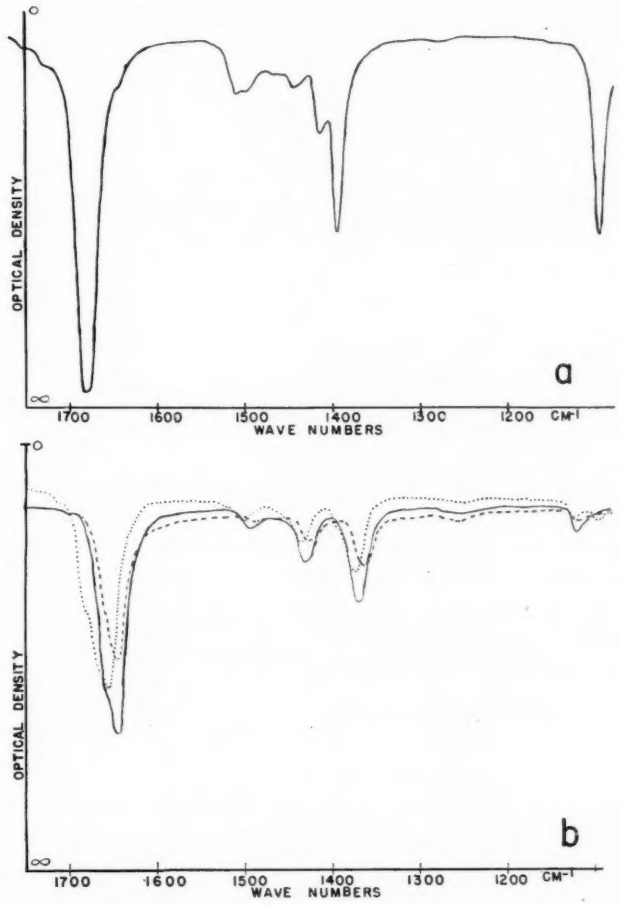


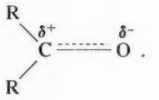
FIG. 1. Infrared spectra in methylene chloride solutions.
(*a*) ——— N,N-dimethyl formamide.
(*b*) ——— $TiCl_4.2HCON(CH_3)_2$; --- $TiBr_4.2HCON(CH_3)_2$; . . . $TiI_4.2HCON(CH_3)_2$.

temperature for the iodide complex. This behavior is also in agreement with the decreasing electronegativities and increasing steric hindrance of the halogens from chlorine to iodine.

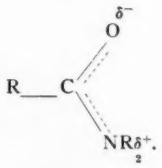
It is noticeable that the infrared spectrum of the iodide complex showed the carbonyl frequency and the CN frequency of the free amide. This is an indication of the presence of some unco-ordinated amide associated with the complex.

The shifts of the carbonyl frequencies in the spectra of these complexes are quite smaller than those observed for other similar complexes. We prepared the complex $TiCl_4.2CH_3COCH_3$ and observed a shift of 47 cm^{-1} when the spectrum was measured in methylene chloride solution. The small shifts observed with amide complexes can be explained by the fact that, in amides, one is dealing with a conjugated system.

In a carbonyl compound like acetone the electronic pull of a Lewis acid is confined to the carbonyl group



In amides, where the carbonyl group is part of a conjugated system, three atoms are involved in the electronic rearrangement caused by co-ordination.



The Complexes of N,N-*Diphenyl Formamide*

The infrared spectra of the complexes of titanium tetrachloride and titanium tetrabromide with N,N-diphenyl formamide are listed in Table II and illustrated in Fig. 2.

TABLE II

Frequencies of N,N-diphenyl complexes (solutions in methylene chloride)

$HCON(C_6H_5)_2$	$TiCl_4.2HCON(C_6H_5)_2$	$TiBr_4.2HCON(C_6H_5)_2$	Assignment
1715 sh			
	1685 w		
1689 vs	1623 vs	1623 vs	ν(C=O)
1602 s	1580 vs	1582 vs	Aromatic ring vibrations
1501 s	1501 s	1501 s	
1456 w	1461 w	1462 w	
1344 m			
1305 w	1373 s	1373 s	CN vibrations
1304 w			
1137 w			

A carbonyl frequency shift of 66 cm^{-1} was measured in the spectra of the halide complexes. Furthermore, a very broad band with four maxima at 1344, 1305, 1304, and 1260 cm^{-1}, which is in the region usually assigned to the CN stretching vibrations, is replaced in the complexes by a narrow band at 1373 cm^{-1}. If one takes 1310 cm^{-1} as the center of the broad band, a positive shift of about 60 cm^{-1} is calculated for the CN vibrations. These shifts are twice as large as those observed in the spectra of similar complexes of N,N-dimethyl formamide. The proximity of the phenyl groups on the nitrogen of the amide seems to increase the basic character of the carbonyl. This is evidenced not only by the increase of the shifts but also by the perturbations of the aromatic ring vibrations.

Susz (11) has mentioned that, for the complexes of benzophenone and acetophenone with Lewis acids, coupling of the carbonyl and phenyl vibrations appears to be appreciable, the phenyl band at 1580 cm^{-1} being considerably enhanced. An increased resonance involving electrons of the carbonyl and the phenyl groups would likely account for this behavior. The electronic pull of a Lewis acid has the same effect as the presence on the

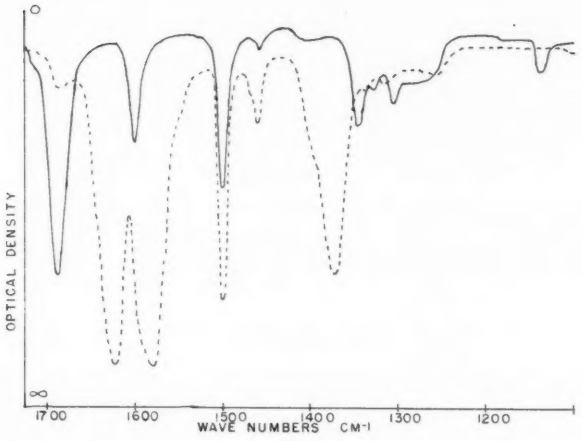


FIG. 2. Infrared spectra in methylene chloride solutions.
——— N,N-diphenyl formamide; – – – $TiCl_4.2HCON(C_6H_5)_2$.

aromatic ring of a nucleophilic parasubstituent. The spectra of N,N-diphenyl formamide complexes exhibit a similar coupling of the carbonyl band with a reinforced phenyl band, appearing at 1580–1582 cm^{-1} instead of 1602 cm^{-1} in the pure amide. In this case, the aromatic ring is not adjacent to the carbonyl, but the electron delocalization induced by the Lewis acid is likely to be transmitted through the conjugated N—C—O skeleton. This statement is further supported by Randall (12) and Bellamy (13), who consider the appearance of a band at 1580 cm^{-1} as an indication of a conjugated phenyl.

ACKNOWLEDGMENTS

The authors are indebted to the National Research Council of Canada for its financial assistance (R. R.) and to l'Office de Recherches Scientifiques de la Province de Québec, for a fellowship to one of us (J. A.). They are also grateful to Prof. Camille Sandorfy for very useful discussions on the infrared spectra measured.

RÉSUMÉ

On a préparé des complexes de coordination des halogénures de titane (IV) avec la N,N-diméthyl formamide et la N,N-diphenyl formamide. Les spectres infrarouges comparés des amides pures et des complexes nous ont permis de prouver que l'effet bathochrome de la fréquence carbonyle est fonction de l'halogène lié au titane (IV) et aussi de la nature des substituants sur l'azote de la molécule de l'amide. Tous les composés préparés sont expliqués en supposant que la coordination au titane se fait par l'oxygène de l'amide.

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THE ABSTRACTION OF HYDROGEN ATOMS FROM AMINES AND RELATED COMPOUNDS[1]

R. K. BRINTON

ABSTRACT

The reaction of methyl radicals with a group of amines and amine-like compounds has been investigated in the temperature range 125° to 157° C. The abstraction activation energies of hydrogen atoms from these compounds, the corresponding pre-exponential factors, and the actual reaction rates indicate that the N—H hydrogen atoms are more labile than the C—H atoms in these compounds.

INTRODUCTION

The abstraction of hydrogen atoms by methyl radicals from various compounds has been studied in detail by Trotman-Dickenson and Steacie (1). In the case of those compounds with hydrogen atoms attached only to the carbon atom skeleton, they were able to correlate the abstraction activation energies, pre-exponential factors, and actual rates of reaction with the structure of the molecules. Compounds such as amines and alcohols having hydrogen atoms attached to nitrogen or oxygen as well as carbon atoms were found more difficult to evaluate, and they made no attempt to assess the effect of structural factors or the relative importance of the N—H and O—H hydrogen as compared with the C—H hydrogen in the over-all contribution to methane formation.

A little more information on this problem was furnished by the work of Brinton and Volman (2), who reported the activation energy for the methyl radical abstraction of hydrogen from ethylenimine to be 4.8 kcal/mole, a value much lower than any C—H type abstraction found by Trotman-Dickenson and Steacie. Brinton and Volman felt that these data indicated a greater lability of N—H hydrogen than C—H hydrogen in this compound. However, their tacit assumption that the N—H hydrogen was the active species in the abstraction was not founded on any experimental evidence.

Recently Kozak and Gesser (3) investigated the photolysis of triethylamine. In the course of the study they determined the activation energies of methane formation in the reaction of methyl radicals with trimethylamine, triethylamine, and diethylamine by a technique similar to that of Trotman-Dickenson and Steacie. The values obtained, 8.0, 5.3, and 5.7 kcal/mole respectively, led them to conclude that the hydrogen attached to carbon was being abstracted in all of these compounds, and that this was probably the case for ethylenimine also.

Confirmation of Kozak and Gesser's mechanism is to be found in the course of a number of condensations between the α-carbon of an amine and a 1-olefin occurring in the liquid phase (4). These reactions are induced by peroxides or light so a radical process may be inferred, and it is probable that an abstraction similar to that postulated by Kozak and Gesser is responsible for the type of products found.

The present study was initiated in an attempt to clarify the kinetics of the reaction of methyl radicals with amines. It was hoped that if a clear difference between the ease of hydrogen abstraction from carbon and nitrogen atoms does indeed exist, sufficient data could be obtained so predictions could be unequivocally made concerning the course of the abstraction reactions for amine-type compounds.

[1]*Manuscript received April 4, 1960.*
Contribution from the Department of Chemistry, University of California, Davis, California.

EXPERIMENTAL

The apparatus and experimental technique have been described previously (5). In the present study the methane and ethane fractions were separated at −215° (solid nitrogen) and −185° (liquid oxygen). Mass spectrometric analyses were made of both fractions. The −185° fraction was in all cases more than 99% ethane. The −215° fraction was 70–99% methane, depending on the sample size. Very small samples (<1 μ mole) contained appreciable amounts of nitrogen produced from the residual air left in the rather large (3 liters) reaction system.

Di-*t*-butyl peroxide was kindly supplied by the Shell Development Co.; it was purified by bulb-to-bulb distillation *in vacuo*. A center fraction was used in the pyrolysis experiments.

Methylamine and ethylamine were generated from the hydrochloride salts with saturated KOH solution. The amines were dried over KOH pellets, degassed, and distilled *in vacuo*.

Dimethylamine, diethylamine, and diisopropylamine were Eastman White Label products. They were used without further purification except for the degassing procedure.

t-Butyl ethylenimine was supplied by Dr. Albert Bottini (6).

t-Butyl acetaldimine was prepared according to the method of Hurwitz (7) from *t*-butylamine and acetaldehyde. The fraction of boiling range 78.0–78.1° (uncorr.) from a 20-plate helice column was collected for the pyrolyses.

The acetaldazine sample used was that employed in previously described work from this laboratory (8).

The pyrolyses were carried out at two temperatures: 156.73° C at which a 4-minute reaction period was used, and 124.82° where 2 to 4 hours were required for sufficient reaction. The DTBP concentration was about the same in all experiments, 4.6×10^{-7} mole cm^{-3} (ca. 11 mm). The pressure of the added compound was adjusted, insofar as it was possible, to produce comparable and significant amounts of methane and ethane so that the analyses of these gases were of good precision. Approximately 3.5% of the DTBP was decomposed at 156.73° and from 2 to 4% at 124.82°. In most trials the percentage of the added compound reacting was from 0.3 to 2%. In two experiments (dimethylamine and diethylamine at 124.82°) the percentage reacted was about 6%. The average concentration of the added compound was used in all calculations.

RESULTS AND DISCUSSION

The pyrolysis of DTBP has been frequently used as a source of methyl radicals in this laboratory (2, 5, 9, 10, 11). The important reactions are:

$$(CH_3)_3COOC(CH_3)_3 \xrightarrow{k_1} 2(CH_3)_3CO \qquad [1]$$

$$(CH_3)_3CO \xrightarrow{k_2} CH_3COCH_3 + CH_3 \qquad [2]$$

$$2CH_3 \xrightarrow{k_3} C_2H_6 \qquad [3]$$

$$CH_3 + (CH_3)_3COOC(CH_3)_3 \xrightarrow{k_4} CH_4 + (CH_3)_3COO(CH_3)_2CH_2 \qquad [4]$$

If another hydrogen-containing source, RH, is present additional methane is formed by the process

$$CH_3 + RH \xrightarrow{k_5} CH_4 + R. \qquad [5]$$

In the present study the total rate of CH_4 production, $R_{CH_4}^{total}$, the summation of [5] and [4], was corrected for the CH_4 formed by [4] in order to evaluate the rate of CH_4 formed in process [5] only, R_{CH_4}. This correction was made by the following formula:

$$R_{CH_4} = R_{CH_4}^{total} - R_{CH_4}^{DTBP}[R_{C_2H_6}/R_{C_2H_6}^{DTBP}]^{1/2}.$$

Here $R_{CH_4}^{DTBP}$ and $R_{C_2H_6}^{DTBP}$ are the rates of CH_4 and C_2H_6 production in a pyrolysis containing only DTBP. The quantity $R_{C_2H_6}$ is the rate of C_2H_6 production in a pyrolysis containing RH as an additive. As the concentration of DTBP was about identical in all trials, the ratio of rates of methane formation by process [4] of any two pyrolyses would be expected to be equal to the half power of the ratio of the corresponding rates of ethane formation in the same two experiments. In most of the experiments the correction calculated from $R_{CH_4}^{DTBP}$ was relatively small, 3–15% of the $R_{total}^{CH_4}$. However, the *t*-butyl acetaldimine reacted so slowly that the corrections in these pyrolyses were approximately equal to the methane formed by the hydrogen abstraction from acetaldimine itself.

The values of $R_{CH_4}/R_{C_2H_6}^{1/2}$ [M] = $k_5/k_3^{1/2}$ have been calculated for the listed concentration, [M], of various amine and amine-like additives. These quantities as well as the activation energies, E_5, and the log A_5 are listed in Table I. Gomer and Kistiakowsky's (12) values of $E_3 = 0$ and $A_3 = 4.5\times10^{13}$ for the combination of methyl radicals were used to calculate the individual E_5 and log A_5 in the table. For most compounds, determinations were made at two temperatures only. The previous studies with ethylenimine (2) showed the Arrhenius plot to be a straight line at least over the rather limited temperature range imposed by the use of DTBP as a methyl radical source. Pyrolyses using dimethylamine and diethylamine were made at 138.6° C as well as for the two temperatures shown in Table I. The values of $k_5/k_3^{1/2}$, 12.05 and 15.31 respectively, gave good straight lines in the Arrhenius plots of these two compounds.

The data of Table I give strong indication that the N—H hydrogen is active in the formation of methane in the methyl radical abstraction process rather than the hydrogen on the carbon alpha to the nitrogen. This evidence is most easily demonstrated by considering certain groups of Table I separately.

TABLE I

The reaction of methyl radicals with amines and related compounds

Reactant (M)	124.8° C		156.73° C		E_5, kcal	log A_5
	[M]×10⁷	$R_{CH_4}/R_{C_2H_6}^{1/2}$[M]	[M]×10⁷	$R_{CH_4}/R_{C_2H_6}^{1/2}$[M]		
Methylamine	2.73	0.944	2.73	1.92	7.6	10.9
Ethylamine	1.02	2.81	1.90	5.45	7.1	11.2
Dimethylamine	0.982	8.77	1.68	17.30	7.2	11.7
Diethylamine	0.97	11.34	1.89	22.20	7.2	11.8
Diisopropylamine	1.00	8.70	0.755	18.02	7.8	12.0
Ethylenimine	3.0	16.6[a]	—	—	4.8	10.7
t-Butyl ethylenimine	8.38	0.14	10.78	0.26	6.6[b]	9.6[b]
t-Butyl acetaldimine	8.62	1.09	10.66	2.26	7.8	11.1
Acetaldazine	5.89	2.26	5.33	3.99	6.1	10.5

NOTE: All rates and concentrations are expressed in units of moles, cm³, and sec.

[a]Data on ethylenimine are from ref. 2.

[b]These calculated values are of low accuracy because of the very low R_{CH_4} and the relatively large correction, $R_{CH_4}^{DTBP}$, used in the R_{CH_4} calculation.

1. Dimethylamine (DMA), *Diethylamine* (DEA), *and Diisopropylamine* (DIA)

These compounds, all secondary amines, have quite different structures of the two

aliphatic side chains. All six of the aliphatic group hydrogens of DMA are on primary carbons, four of the 10 hydrogens of DEA are on secondary carbons, and two of the 14 hydrogens of DIA are on tertiary carbons. Table I indicates that all three of these compounds have very similar abstraction activation energies, E_5, and pre-exponential factors A_5. Trotman-Dickenson and Steacie (1) and Trotman-Dickenson (13) found the analogous reactions in hydrocarbon systems to be quite sensitive to the type of carbon containing the abstractable hydrogen. If the side-chain hydrogens were being abstracted from the amines, it would be expected that the three amines being considered would show a similar trend, say to the analogous series ethane, *n*-butane, and 2,3-dimethyl butane, which have respectively the abstraction energies 10.4, 8.3, and 6.9 kcal/mole (1). In addition, if DMA is assumed to have six abstractable hydrogen atoms (primary), DEA four (secondary), and DIA two (tertiary) the actual values of $R_{CH_4}/R_{C_2H_6}$ [M] at 124.8° per "active" hydrogen are in the ratios of 1:1.9:3.0. The corresponding ratios for ethane, *n*-butane, and 2,3-dimethyl butane found by Trotman-Dickenson and Steacie (1) are 1:7.5:51.7 at 182° C. Thus neither the abstraction activation energies nor the rates of abstraction indicate that the side-chain hydrogens of the amines are participating in the abstraction process. It must be inferred rather that the amine hydrogen is active in the reaction and the similarity of activation energy, pre-exponential factors, and the actual rates of abstraction may be attributed to their being secondary amines.

2. Ethylenimine and t-Butyl Ethylenimine

The abstraction activation energy for ethylenimine given previously (2) is considerably lower than others reported for H-atom abstraction by methyl radical from organic compounds. In this former investigation it was assumed that the N—H bond was the active participant in the reaction. It is proposed in the present study to test the validity of such a mechanism by reacting the methyl radicals with *t*-butyl ethylenimine wherein the N—H hydrogen has been replaced by the *t*-butyl group. Table I shows the rate of abstraction from the *t*-butyl ethylenimine to be less than 10% of the analogous reaction from ethylenimine itself at 124.8°. Inasmuch as the ring hydrogens are available to an equal degree from both these compounds, a strong argument can be made to eliminate these H atoms as sources of methane production and to establish the N—H hydrogen as the principal reactant with the methyl radicals. The activation energy and pre-exponential factor for *t*-butyl ethylenimine are shown bracketed in Table I. The very low rate of methane production for this compound at both experimental temperatures introduces considerable error into these calculated quantities. They should be considered as approximations only and no significance can be attributed to their values relative to the values of the corresponding constants for ethylenimine.

3. t-Butyl Acetaldimine and Acetaldazine

t-Butyl acetaldimine $(CH_3)_3C—N{=}CH—CH_3$ and acetaldazine $CH_3CH{=}N—N{=}CHCH_3$ have been included to show the behavior of compounds which do not have N—H hydrogens available. Again it is not certain which hydrogen is entering into reaction, those of the terminal CH_3 group or the H atoms on the carbon of N=C bond. However, it is quite significant that the rate at 124.8° for the acetaldazine abstraction is almost exactly twice that for *t*-butyl acetaldimine. Two hydrogens on the C=N carbon and six hydrogens on CH_3 groups are available in acetaldazine, and one hydrogen on C=N carbon and 12 hydrogens on CH_3 groups are contained on the *t*-butyl acetaldimine. Thus the

abstraction rates of the two compounds are compatible with the interpretation that the C=N carbon atom furnishes the hydrogen for the abstraction although the evidence is not extensive enough to justify a definite conclusion.

4. Methylamine and Ethylamine

The data in Table I for these two compounds do not furnish further confirmation for the thesis that the N—H hydrogen is active in the abstraction process. On the other hand the evidence is not necessarily contradictory. The similarity in the abstraction activation energies does tend to support the idea of N—H hydrogen abstraction but it is not clear why the abstraction rate at 124.8° for ethylamine is about three times that of methylamine.

The conclusions reached in the preceding discussion as well as some of the experimentally determined data differ markedly from those of Kozak and Gesser's study (3). Their conclusion that low and similar abstraction energies for triethylamine and diethylamine probably indicate a labile C—H rather than N—H hydrogen is reasonable on basis of this evidence alone. However, their value for the abstraction constant of diethylamine ($\log A = 10.9$,* $E = 5.7$) is somewhat different from that shown in Table I ($\log A = 11.8$, $E = 7.2$). Reasons for this difference are speculative as the experimental methods used in both investigations have proved to be reliable in a large number of previously tested cases. It should be pointed out that the data of Trotman-Dickenson (13) for methylamine ($\log A = 11.3$, $E = 8.4$) and dimethylamine ($\log A = 11.3$, $E = 7.2$) are in good agreement with the corresponding quantities shown in Table I, and in addition their value for trimethylamine ($\log A = 11.8$, $E = 8.8$) agrees reasonably well with Kozak and Gesser's result ($\log A = 12.6$, $E = 8.0$) for this same compound. Thus all three studies show conformity between at least some compounds that were mutually tested.

A possible clue to the wide difference in abstraction activation energies for diethylamine between the present study and Kozak and Gesser's work may be seen in their Arrhenius plots for diethylamine and triethylamine. In both curves there is a pronounced falling off in the values of $k_{abs}/k_{comb}^{1/2}$ at the higher temperature. Their explanation for this curvature, the growing importance of the reaction,

$$CH_3 + (C_2H_5)_2NH \rightarrow C_2H_5NHCH_2 + C_2H_6, \qquad [6]$$

as temperature is increased is compatible with the experimental evidence. However, their assumption that the straight lines drawn through the limited low-temperature points truly represent the behavior of reaction [5] only is open to considerable doubt. It seems reasonable that the effect of reaction [6] is active to a degree, at least, over the whole temperature range and that the slopes of the Arrhenius lines should be somewhat steeper, thereby yielding higher activation energies for triethylamine and diethylamine. Whether both curves would be affected in a similar fashion is not predictable, but it is not unlikely that the relative magnitude of the two activation energies might shift so that different conclusions could be drawn concerning the role of the various hydrogens in the abstraction process.

The argument in the previous paragraph would apply equally as well to the present study if reaction [6] were taking place even to a limited extent in the lower temperature range. As the total rate of ethane production would then be the sum of processes [3] and

**The values calculated by Kozak and Gesser for this corresponding quantity for diethylamine, $A_{19}/A_8^{1/2} = 9\times10^{-5}$ is in error. The value should be 1.4×10^{-8} molecule$^{-1/2}$ cm$^{3/2}$ sec$^{-1/2}$. Similarly their value of $A_{11}/A_8^{1/2}$ for triethylamine shown as 3×10^{-5} should be 6.8×10^{-9}.*

[6], the error made by using the total $R_{C_2H_6}$ in place of $R_{C_2H_6}$ by [3] would depend on the relative rates of the two ethane-producing reactions. It can be shown that

$$\frac{R_{CH_4}}{R_{C_2H_6}{}^{1/2}[M]} = \frac{k_5}{k_3{}^{1/2}}\left[1+\frac{k_6[M]}{k_3[CH_3]}\right]^{-1/2}.$$

Thus the effect of process [6] on the value determined for $k_5/k_3{}^{1/2}$ will be minimized by low concentrations of additive, [M], as well as high methyl radical concentrations, $[CH_3]$.

It is significant that there is appreciable difference between the magnitude of both these variables in the two studies under comparison. The diethylamine concentrations used by Kozak and Gesser in their photolyses (12.4×10^{-7} to 16.9×10^{-7} mole cm^{-3}) were about 10 times those used in the DTBP pyrolyses. Clearly in this case the effect of the additive term in the above expression will be only about 1/10th as great for this study as for Kozak and Gesser's work.

The consequence of using an appreciably higher methyl radical steady-state concentration at 156.7° in the peroxide work than in the photolyses is not so easily evaluated. In the present work on diethylamine $R_{C_2H_6}$ is approximately 3.0×10^{-12} and 115×10^{-12} mole cm^{-3} sec^{-1} at 124.8° and 156.7° respectively. The lower temperature rate is about comparable to that of Kozak and Gesser's 123° experiment (2.5×10^{-12}), but the higher temperature rate is about 50 times the rate of their 155° photolysis (2.1×10^{-12}). From the above expression it may be seen that for a constant [M] the activation energy measured by the change in $R_{CH_4}/R_{C_2H_6}{}^{1/2}$ [M] as a function of temperature will be exactly that of E_5 if $k_6/[CH_3]$ is maintained constant over the range of temperature variation. Kozak and Gesser's Arrhenius curves for diethylamine and triethylamine indicate that E_6 is considerably higher than E_5, perhaps of the order of 15 kcal/mole, and therefore $k_6(156.7°)/k_6(124.8°) \cong 4$. In order for $k_6/[CH_3]$ to maintain its constancy the ratio $R_{C_2H_6}(156.7°)/R_{C_2H_6}(124.8°)$ must be of the order of 16. If this estimate of E_6 is valid and if $R_{C_2H_6}{}^{1/2}$ may be taken as being proportional to $[CH_3]$, the experimental technique of Kozak and Gesser in which $[CH_3]$ was about constant would yield a value somewhat too small for E_5 while that of the present work would give too large an activation energy. The magnitude of the error of course depends on the size of the additive term within the bracketed expression of the above formula. Determination of the absolute magnitude of the effect of [6] on ethane formation and thus on the value of E_5 is perhaps possible by a detailed study of diethylamine variation at constant temperature in the case of the DTBP pyrolyses. In a like manner investigation of the effect of intensity and diethylamine variation at constant temperature for the photolyses of Kozak and Gesser should produce similar information.

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INTERACTION OF TWO PARALLEL PLATES IN A BINARY ASYMMETRICAL ELECTROLYTE

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ABSTRACT

The interaction of two charged plate-like parallel colloidal particles immersed in a binary asymmetrical electrolyte of valency type $1-m$, where the coagulating ion is univalent, is considered. Applying the Derjaguin–Verwey–Overbeek theory of colloid stability, it is assumed that the plate surface potential is independent of the plate separation. The Stern modification in the absence of specific adsorption of the coagulating ion is also considered. It is demonstrated that as the valency m of the non-coagulating ion increases, the concentration of the coagulating ion at flocculation increases somewhat, but not to the extent predicted by Ostwald's activity coefficient rule. Application of the Stern theory reduces this trend and may even produce a decrease, with increasing m, in the coagulating ion concentration, particularly for a 1–2 electrolyte.

1. INTRODUCTION

According to the Schulze–Hardy rule, the concentration of a binary electrolyte required to flocculate a hydrophobic sol depends mainly on the valency of the ion whose charge is opposite in sign to that on the particles and only slightly on the valency of the ion with a charge of the same sign as that on the particles. The stability theory of Derjaguin and Landau (1) and Verwey and Overbeek (2) reproduces approximately the experimental dependence of the limiting concentration for coagulation on the valency of the flocculating (oppositely charged) ion. However, no systematic theoretical investigation of the corresponding (secondary) dependence on the valency of the like-charged ion has been made and it is the purpose of this paper to determine the relationship which is derived from the above theory for the case of a univalent flocculating ion. According to Ostwald and his collaborators (3), the coagulation concentrations for a number of hydrophobic sols are described approximately by Ostwald's empirical rule, which states that the electrolyte in the dispersion medium flocculates at the same value for the coefficient of activity of the dominating (coagulating) counter-ion. If the Debye–Huckel theory of strong electrolytes at dilute concentrations is applied to a binary electrolyte whose ion types 1 and 2 have densities n_1^0 and n_2^0 and valencies z_1 and z_2 respectively, then Ostwald's rule states that for a given colloidal system, the activity coefficient f_1 of the coagulating ion (which is assumed to be of type 1) is given by

$$[1.1] \qquad \frac{1}{\sqrt{m}} \log f_1 = -z_1^2 \left[\frac{\pi e^6}{(\epsilon kT)^3} \frac{(n_1^0 z_1^2 + n_2^0 z_2^2)}{m} \right]^{\frac{1}{2}} = \text{constant},$$

where e is the electronic charge, ϵ the dielectric constant of the electrolyte medium, k Boltzmann's constant, T the absolute temperature, and m the number of ions of type 1 in a molecule. For a univalent coagulating ion

$$[1.2] \qquad m = -(z_2/z_1) = n_1^0/n_2^0,$$

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making use of the condition of electrical neutrality

$$n_1^0 z_1 + n_2^0 z_2 = 0. \qquad [1.3]$$

The factor $1/\sqrt{m}$ in [1.1] does not occur in the theory of electrolytes so that Ostwald does not employ the usual definition of activity coefficient. The introduction of this factor in colloids emphasizes the significance of the coagulating ion concentration. It follows from Ostwald's rule that for an electrolyte series of valency types $z_1 - z_2$ where $z_1 = 1$ and z_2 takes the values $-1, -2, -3, \ldots$, the quantity $n_1^0(1+m)/m$ should be a constant at the flocculating electrolyte concentration. Various objections, however, have been made to Ostwald's activity coefficient rule. Weiser (4) has commented on the wide spread in the experimental results over which Ostwald averaged and Weiser and Milligan (5) have demonstrated its limited applicability. Tezak and his school (6, 7) maintain that most of the activity coefficients compiled by Ostwald refer to concentrated sols, whereas a physical meaning to the activity coefficient of the flocculating ion can only be given for small sol concentrations. In view of the different premises on which the current theory of colloid stability is based, the Ostwald rule is inadequate theoretically and should be regarded as an empirical result.

In this paper, the interaction of two parallel plate-like particles immersed in a large volume of a binary electrolyte will be examined. It is assumed that the plates are negatively charged, the coagulating cation is univalent ($z_1 = 1$), and the valency z_2 of the anion may have the values $-1, -2, -3, -4$, or $-\infty$. Following Verwey and Overbeek (2), the charge on the plates is supposed to be due to potential-determining ions whose concentration in the dispersion medium is negligible and the binary flocculating electrolyte is indifferent, i.e. there is no specific adsorption of the cation on the plate surface. This means that the potential at the plates is constant, independent of plate separation and of both the nature and concentration of the coagulating electrolyte. Extensive work on the interaction of two parallel plates in a binary symmetrical electrolyte have been carried out by a number of authors (1, 2, 8). The corresponding problem in a 1–2 electrolyte has been treated by Levine and Suddaby (9) and related work on the solution of the Poisson–Boltzmann equation in electrolyte mixtures containing ions of valencies 1 to 4 has been published by Wood and Robinson (10, 11, 12). Electrolyte mixtures of univalent and divalent ions have also been investigated by Glazman and his collaborators (13, 14). Provided the electrolyte in the dispersion medium contains only univalent and divalent ions, it is possible to express the interaction energy of the electric double layers of the plates in terms of elliptic integrals which are tabulated. However, with electrolytes containing higher valency ions, this interaction is obtained in terms of hyperelliptic integrals which are not tabulated. We shall avoid the use of such integrals by expanding the expression for the interaction energy in the form of a series which converges rapidly provided the plate separation is not too large and the surface potential not too small. This will permit the determination of the conditions for flocculation. The investigations of the various authors mentioned above are all based on the Gouy–Chapman theory, in which the diffuse layer is assumed to extend right up to the plate walls. However, we shall also consider the modification (due to Stern) which corrects for the finite size of the ions in the absence of specific adsorption.

By an obvious change in the sign of the surface potential the formulae derived in this paper can be applied to a $m-1$ valency electrolyte, in which the coagulating ion has the valency m. This will be considered in a later paper.

2. GENERAL EXPRESSION FOR INTERACTION ENERGY

Let $2h$ be the separation of the plates and x the distance measured from the median plane. If the linear dimensions of the plates are large compared with both h and the so-called thickness of the diffuse layers, then the electrostatic potential ψ at any position x will be a function of x only. It is assumed that the potential $\psi = \psi(x)$ satisfies the classical Poisson–Boltzmann equation throughout the range $-h < x < h$ of the overlapping diffuse layers in the region between the plates. This equation reads

$$\frac{d^2\psi}{dx^2} = -\frac{4\pi e}{\epsilon}[n_1^0 z_1 \exp(-\beta z_1 e\psi) + n_2^0 z_2 \exp(-\beta z_2 e\psi)], \qquad [2.1]$$

where $\beta = 1/kT$. The potential at the median plane will be $\psi_m = \psi(0) < 0$ and that at each plate $\psi_0 = \psi(h) < \psi_m$, the potential zero being taken in the interior of the electrolyte dispersion medium. The classical electrostatic energy per unit area of the two plates is given by

$$E_e(h) = \frac{\epsilon}{4\pi}\int_0^h \left(\frac{d\psi}{dx}\right)^2 dx. \qquad [2.2]$$

The ideal (osmotic) pressure due to the thermal motion of the ions at any x in the diffuse layers where the local ion densities are n_1 and n_2 and the corresponding pressure in the electrolyte interior are respectively,

$$\Pi = kT(n_1+n_2) \quad \text{and} \quad \Pi_0 = kT(n_1^0+n_2^0). \qquad [2.3]$$

Also the surface charge density on each plate is

$$\sigma = \frac{\epsilon}{4\pi}\left(\frac{d\psi}{dx}\right)_{x=h}. \qquad [2.4]$$

There are a number of completely equivalent expressions for the free energy to be associated with the electric double layers of the two plates. Thus, per unit area of plates, this energy can be written as

$$F(h) = -2\int_0^{\psi_0} \sigma d\Psi = -E_e(h) - 2\int_0^h (\Pi - \Pi_0)dx. \qquad [2.5]$$

These forms include the so-called electrical and chemical contributions to the free energy and were first used by Verwey and Overbeek (2); a general derivation has been given by Levine (15). In the first integral in [2.5] it is imagined that the surface charge density σ varies from 0 to its full (actual) value σ_0 and the surface potential is Ψ which is a function $\Psi(\sigma)$ of σ; thus $\psi_0 = \Psi(\sigma_0)$. The corresponding double layer force of repulsion per unit area (the electric double layer pressure) can be expressed in the forms

$$p_e(h) = -\frac{1}{2}\left(\frac{\partial F}{\partial h}\right)_{\psi_0} = kT(n_1^{(m)}+n_2^{(m)}-n_1^{(0)}-n_2^{(0)}), \qquad [2.6]$$

where the derivative with respect to h is taken at constant plate potential ψ_0 and $n_1^{(m)}$ and $n_2^{(m)}$ are the ion densities at the median plane $x = 0$. The first form was derived by Verwey and Overbeek (2) and the second independently by Frumkin and Gorodetskaja (16) and by Langmuir (17). A first integral of the Poisson–Boltzmann equation [2.1] is

$$\frac{d\psi}{dx} = -\left[\left(\frac{8\pi kT}{\epsilon}\right)\sum_{i=1}^{2} n_i^0 \{\exp(-\beta z_i e\psi) - \exp(-\beta z_i e\psi_m)\}\right]^{\frac{1}{2}} \qquad [2.7]$$

for $x > 0$, remembering that the above derivative is negative when the plates are negatively charged. This may also be expressed as

$$\Pi - \Pi_0 - \frac{\epsilon}{8\pi}\left(\frac{d\psi}{dx}\right)^2 = p_e(h), \tag{2.8}$$

where the left-hand side refers to any point x in the diffuse layers. It follows from [2.5] and [2.8] that

$$F(h) = -2E_e(h) - 2hp_e(h). \tag{2.9}$$

In addition to the double layer repulsion the van der Waals attraction between the plates must be considered. If the thickness of the plates is large compared with h, the energy of attraction per unit area takes the particularly simple form (1, 2)

$$W(h) = -A/48\pi h^2, \tag{2.10}$$

where the constant A depends upon the material of the plates and of the dispersion medium in which the plates are immersed. The resultant interaction energy per unit area of the two plates is

$$U(h) = F(h) + W(h) - F(\infty). \tag{2.11}$$

Following Derjaguin and Landau (1) and Verwey and Overbeek (2) we shall use the following conditions to define the transition from the stable to unstable sol (i.e. the critical flocculating relations)

$$U(h) = 0, \qquad [\partial U(h)/\partial h]_{\psi_0} = 0. \tag{2.12}$$

The second condition in [2.12] may be expressed as

$$p_e(h) = A/48\pi h^3. \tag{2.13}$$

Making use of [2.10] and [2.13], the first relation in [2.12] can be put in the form

$$2E_e(h) + 3hp_e(h) + F(\infty) = 0. \tag{2.14}$$

3. GENERAL FORM OF COAGULATION CONDITIONS

It is convenient to introduce a Debye–Huckel parameter K for the positive ions only, defined by

$$K^2 = 8\pi n_1^0 z_1^2 e^2/\epsilon kT, \tag{3.1}$$

and also the quantities

$$\xi = \exp(\beta z_1 e\psi_m) < 1, \qquad u = \xi \exp(-\beta z_1 e\psi) > 1 \tag{3.2}$$

and

$$f(u, \xi, m) = \left[\exp(-\beta z_1 e\psi) - \exp(-\beta z_1 e\psi_m) + \frac{1}{m}\left\{\exp(-\beta z_2 e\psi) - \exp(-\beta z_2 e\psi_m)\right\}\right]^{\frac{1}{2}} \tag{3.3}$$

$$= \xi^{-\frac{1}{2}}\left[u - 1 + \frac{1}{m}\xi^{m+1}(u^{-m} - 1)\right]^{\frac{1}{2}}.$$

Then [2.7] becomes

$$\frac{d\psi}{dx} = -\left(\frac{8\pi n_1^0 kT}{\epsilon}\right)^{\frac{1}{2}} f(u, \xi, m), \tag{3.4}$$

so that the formula [2.2] for the electrostatic energy reads

$$[3.5] \qquad E_e(h) = \frac{\epsilon}{4\pi}\int_{\psi_m}^{\psi_0}\frac{d\psi}{dx}d\psi = \frac{2n_1^0kT}{K}\int_1^{u_0}\frac{f(u,\xi,m)}{u}du,$$

where $u_0 = \xi\exp(-\beta z_1 e\psi_0)$. Also the double layer pressure [2.6] is

$$[3.6] \qquad p_e(h) = n_1^0kT\left(\frac{1}{m}\xi^m+\xi^{-1}-\frac{m+1}{m}\right),$$

and from [3.4]

$$[3.7] \qquad Kh = K\int_{\psi_m}^{\psi_0}\frac{dx}{d\psi}d\psi = \int_1^{u_0}\frac{du}{uf(u,\xi,m)},$$

which, indeed, becomes the solution of the Poisson–Boltzmann equation if h and u_0 are replaced by x and u respectively.

When $h = \infty$, $\psi_m = 0$ and therefore $\xi = 1$. To calculate $F(\infty)$ we apply the first relation in [2.5] to unit area of both sides of a single plate immersed in an infinite volume of the electrolyte. The surface charge density at a single plate surface, whose potential is Ψ, is given by

$$[3.8] \qquad \sigma(\infty) = \frac{\epsilon}{4\pi}\lim_{h\to\infty}\frac{d\psi}{dx} = -\left(\frac{n_1^0kT\epsilon}{2\pi}\right)^{\frac{1}{2}}f(v,1,m),$$

where $v = \exp(-\beta z_1 e\Psi)$. Noting that the surface potential ψ_0 is independent of h, it follows that

$$[3.9] \qquad F(\infty) = -2\int_0^{\psi_0}\sigma(\infty)d\Psi = -\frac{8n_1^0kT}{K}J(\psi_0,m),$$

$$[3.10] \qquad J(\psi_0,m) = -\frac{1}{2}\frac{z_1e}{kT}\int_0^{\psi_0}f(v,1,m)d\psi = \frac{1}{2}\int_1^{v_0}\frac{f(v,1,m)}{v}dv,$$

where $v_0 = \exp(-\beta z_1 e\psi_0)$.

Substituting [3.5], [3.6], [3.7], and [3.9] into the coagulation criterion [2.14] and dividing through by $8n_1^0kT/K$ this criterion appears in the form

$$[3.11] \qquad \frac{1}{2}\int_1^{u_0}\frac{f(u,\xi,m)}{u}du+\frac{3}{8}\left(\frac{1}{m}\xi^m+\xi^{-1}-\frac{m+1}{m}\right)\int_1^{u_0}\frac{du}{uf(u,\xi,m)}-J(\psi_0,m) = 0,$$

which, indeed, applies to any binary electrolyte. It is immediately obvious that the electrolyte concentration variable has been eliminated from [3.11] and that for given ψ_0 and m the value of ξ at which coagulation occurs is obtained. Once [3.11] has been solved for ξ, equation [3.7] then gives Kh and the critical value of K can be obtained from [2.13] and [3.6], from which we derive

$$[3.12] \qquad K = \frac{6\epsilon}{A}\left(\frac{kT}{z_1e}\right)^2(Kh)^3\left[\frac{1}{m}\xi^m+\xi^{-1}-\frac{m+1}{m}\right].$$

Certain points about the behavior of the left-hand side of [3.11], as a function of ξ, appear on considering the left-hand side of [2.14], which is the same expression multiplied by $8n_1^0kT/K$. As a preliminary we observe that if $\xi = 1$ (which means $h = \infty$), the integral in [3.5] becomes identical with $2J(\psi_0,m)$. This shows that $F(\infty) = -2E_e(\infty)$

and also, using [2.9], that $\lim_{h\to\infty} hp_e(h) = 0$. Hence, the right-hand side of [2.14] vanishes at $h = \infty$ and so $\xi = 1$ is a solution of [3.11]. Now

$$[3.13] \quad 2E_e(h)+3hp_e(h)+F(\infty) = hp_e(h)-F(h)+F(\infty) = hp_e(h)-2\int_h^\infty p_e(h)dh,$$

making use of [2.9] and the first relation in [2.6]. For large h this quantity is positive since it can be shown (see Levine and Suddaby (8)) that $p_e(h)$ diminishes exponentially with h as h increases. At $h = 0$, it is equal to $F(\infty)$ since both $F(h)$ and $hp_e(h)$ vanish. (The charge density σ vanishes at $h = 0$ if the surface potential ψ_0 is constant.) But $F(\infty)$ is negative and so the expression [3.13] must have a second zero for some finite h. As h increases from 0 to ∞, ξ increases from u_0^{-1} to 1. It follows that there is (at least) one root of [3.11] where $\xi < 1$ (see Fig. 1). This is the required solution of [3.12].

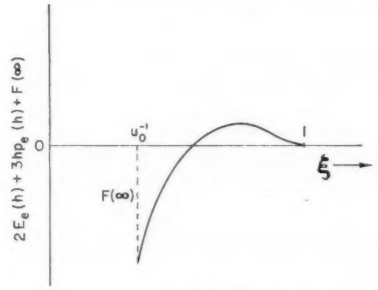

FIG. 1.

4. DETERMINATION OF $F(\infty)$

The expression [3.10] for $J(\psi_0, m)$ can be obtained in closed form for $m = 1, 2, 3, 4$, and ∞. It is helpful to take out a factor $(v-1)v^{-m/2}$ from $f(v, 1, m)$ and then [3.10] becomes

$$[4.1] \quad J(\psi_0, m) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)}{v^{m/2+1}}\left\{v^{m-1}+\frac{(m-1)}{m}v^{m-2}+\ldots+\frac{2v}{m}+\frac{1}{m}\right\}^{\frac{1}{2}} dv,$$

which is integrated as follows for the above values of m.

(i) $m = 1$

$$[4.2] \quad J(\psi_0, 1) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)}{v^{3/2}}dv = v_0^{\frac{1}{2}}+v_0^{-\frac{1}{2}}-2.$$

(ii) $m = 2$

$$[4.3] \quad J(\psi_0, 2) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)}{v^2}(v+\tfrac{1}{2})^{\frac{1}{2}}dv = \frac{1}{2\sqrt{2}}\left\{\frac{(2v_0+1)^{3/2}}{v_0}-3\sqrt{3}\right\}.$$

(iii) $m = 3$

$$[4.4] \quad J(\psi_0, 3) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)}{v^3}(v^3+\tfrac{2}{3}v^2+\tfrac{1}{3}v)^{\frac{1}{2}}dv$$

$$= \frac{1}{3}\left\{\frac{(v_0^3+\frac{2}{3}v_0^2+\frac{1}{3}v_0)^{\frac{1}{2}}}{v_0^2}-\sqrt{2}\right\}+\frac{1}{2}\int_1^{v_0}\frac{(v+\frac{1}{9}v^{-1})dv}{(v^3+\frac{2}{3}v^2+\frac{1}{3}v^3)^{\frac{1}{2}}}.$$

The integrand contains the square root of a cubic in the denominator and so the integral can be expressed in terms of elliptic integrals. By making the substitution

$$cn(t, k) = \frac{v\sqrt{3}-1}{v\sqrt{3}-1}, \qquad k^2 = \frac{\sqrt{3}-1}{2\sqrt{3}},$$

where k is the parameter of the Jacobian elliptic function $cn(t, k)$ (see Byrd and Friedman (18)), integration can be performed. If we put

$$\cos\phi_0 = \frac{v_0\sqrt{3}-1}{v_0\sqrt{3}+1}, \qquad \cos\phi_1 = \frac{\sqrt{3}-1}{\sqrt{3}+1},$$

then it follows that

$$[4.5] \qquad \frac{1}{2}\int_1^{v_0} \frac{(v+\frac{1}{2}v^{-1})dv}{(v^3+\frac{2}{3}v^2+\frac{1}{3}v)^{\frac{1}{2}}} = \frac{2}{3^{5/4}}\left[2E(\phi, k) - F(\phi, k) + \frac{(1+2\cos\phi)(1-k^2\sin^2\phi)^{\frac{1}{2}}}{\sin\phi}\right]_{\phi_1}^{\phi_0}.$$

Here $F(\phi, k)$ and $E(\phi, k)$ are elliptic integrals of the first and second kinds respectively. Substitution of [4.5] into [4.4] gives the required value of $J(\psi_0, 3)$.

(iv) $m = 4$

$$[4.6] \qquad J(\psi_0, 4) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)}{v^3}(v^3+\tfrac{3}{4}v^2+\tfrac{1}{2}v+\tfrac{1}{4})^{\frac{1}{2}}dv$$

$$= -\frac{1}{4}\left[\left\{\frac{(v_0-\frac{1}{2})}{v_0^2}(v_0^3+\tfrac{3}{4}v_0^2+\tfrac{1}{2}v_0+\tfrac{1}{4})^{\frac{1}{2}}\right\} + \frac{5}{8}\int_1^{v_0} \frac{vdv}{(v^3+\frac{3}{4}v^2+\frac{1}{2}v+\frac{1}{4})^{\frac{1}{2}}}\right].$$

The integration may be performed by elliptic substitution as in the case $m = 3$. If the elliptic parameter k equals 0.905943 while

$$\cos\phi_0 = \frac{v_0-0.226242}{v_0+1.437902}, \qquad \cos\phi_1 = \frac{1-0.226242}{1+1.437902} = 0.317387,$$

it turns out that

$$[4.7] \qquad \frac{5}{8}\int_1^{v_0} \frac{vdv}{(v^3+\frac{3}{4}v^2+\frac{1}{2}v+\frac{1}{4})^{\frac{1}{2}}} = 1.140225\left[E(\phi, k) - 0.135951F(\phi, k) + \frac{(1+\cos\phi)(1-k^2\sin^2\phi)^{\frac{1}{2}}}{\sin\phi}\right]_{\phi_1}^{\phi_0}.$$

Substitution of [4.7] into [4.6] yields $J(\psi_0, 4)$.

For higher values of m the integrals will be of the hyperelliptic type, which in general are not tabulated. For $m = 5$ and $m = 6$, for instance, the integrands will involve the square root of a quintic. It is possible to express $J(\psi_0, m)$ as a series, which will be indicated in the following section, but no calculations for $m = 5$ or 6 will be made in this paper.

(v) $m = \infty$

$$[4.8] \qquad J(\psi_0, \infty) = \frac{1}{2}\int_1^{v_0} \frac{(v-1)^{\frac{1}{2}}}{v}dv = (v_0-1)^{\frac{1}{2}} - \tan^{-1}(v_0-1)^{\frac{1}{2}}.$$

In Table I, values of $J(\psi_0, m)$ are given for the two values $\beta z_1 e\psi_0 = -2$ and -4, corresponding to $\psi_0 = 50$ and -100 mv respectively at room temperature, when $z_1 = 1$.

TABLE I

Values of $J(\psi_0, m)$

ψ_0	m	1	2	3	4	∞
−50 mv	Exact	1.086	1.162	1.203	1.228	1.333
	Equation [5.19]			1.216	1.238	1.333
−100 mv	Exact	5.524	5.654	5.714	5.75	5.886
	Equation [5.19]			5.729	5.76	5.886

5. CALCULATION OF COAGULATION VALUES OF ELECTROLYTE

In order to determine the critical value of ξ it is necessary to evaluate the two integrals in [3.11] which come from the equations [3.5] and [3.7] for $E_e(h)$ and Kh respectively. If the substitution $u = 1+y^2$ is made, then

$$[5.1] \qquad \frac{1}{2}\int_1^{u_0}\frac{f(u,\xi,m)}{u}du = \xi^{-\frac{1}{2}}\int_0^{y_0}\frac{y^2}{(1+y^2)}\left[1-\frac{1}{m}\xi^{m+1}\frac{[1-(1+y^2)^{-m}]}{y^2}\right]^{\frac{1}{2}}dy$$

where $u_0 = 1+y_0^2$. Now it is seen from [3.2] that provided the potential at the median plane ψ_m is appreciable, then ξ will be significantly smaller than unity particularly if h is not too large and the surface potential ψ_0 large enough. Since

$$\frac{\xi^{m+1}}{m}\cdot\frac{(1-u^{-m})}{(u-1)} = \frac{\xi^{m+1}}{m}u^{-m}(1+u+\ldots+u^{m-1}) < \xi^{m+1},$$

this suggests that if $\xi < 1$ the integrand in [5.1] be expanded in powers of ξ^{m+1}. The series for the symmetrical case $m = 1$ will be the most slowly converging and we shall retain four terms in this case, namely,

$$[5.2] \qquad \frac{1}{2}\int_1^{u_0}\frac{f(u,\xi,1)}{u}du = \xi^{-\frac{1}{2}}[\tan\theta-\theta-\tfrac{1}{2}\xi^2(\theta-I_2)-\tfrac{1}{8}\xi^4(I_2-I_4)-\tfrac{1}{16}\xi^6(I_4-I_6)+\ldots],$$

where it is convenient to introduce

$$[5.3] \qquad y_0 = \tan\theta, \qquad \xi = u_0/v_0 = \sec^2\theta/v_0,$$

and

$$[5.4] \qquad I_n = \int_0^{y_0}\frac{dy}{(1+y^2)^{n/2+1}} = \int_0^{\theta}\cos^n\alpha d\alpha.$$

This is readily evaluated from the relations

$$[5.5] \qquad I_2 = \tfrac{1}{2}(\theta+\sin\theta\cos\theta), \qquad I_n = \frac{(n-1)}{n}I_{n-2}+\frac{1}{n}\sin\theta\cos^{(n-1)}\theta.$$

Retaining three terms for $m = 2$,

$$[5.6] \qquad \frac{1}{2}\int_1^{u_0}\frac{f(u,\xi,2)}{u}du = \xi^{-\frac{1}{2}}[\tan\theta-\theta-\tfrac{1}{4}\xi^3(\theta-I_4)-\tfrac{1}{32}\xi^6(I_2+I_4-I_6-I_8)+\ldots].$$

For $m \geqslant 3$, it is sufficient to retain only two terms, i.e.

$$[5.7] \qquad \frac{1}{2}\int_1^{u_0}\frac{f(u,\xi,m)}{u}du = \xi^{-\frac{1}{2}}\left[\tan\theta-\theta-\frac{1}{2m}\xi^{m+1}(\theta-I_{2m})+\ldots\right].$$

The integral in [3.7] may be similarly approximated. Thus

$$[5.8] \qquad \int_1^{u_0} \frac{du}{uf(u,\xi,m)} = 2\xi^{\frac{1}{2}} \int_0^{y_0} \frac{1}{(1+y^2)} \left[1-\frac{1}{m}\xi^{m+1}\frac{[1-(1+y^2)^{-m}]}{y^2}\right]^{-\frac{1}{2}} dy,$$

and expanding again, in powers of ξ^{m+1},

$$[5.9] \qquad \int_1^{u_0} \frac{du}{uf(u,\xi,1)} = 2\xi^{\frac{1}{2}}(\theta+\tfrac{1}{2}\xi^2 I_2+\tfrac{3}{8}\xi^4 I_4+\tfrac{5}{16}\xi^6 I_6+\ldots),$$

$$[5.10] \qquad \int_1^{u_0} \frac{du}{uf(u,\xi,2)} = 2\xi^{\frac{1}{2}}\{\theta+\tfrac{1}{4}\xi^3(I_2+I_4)+\tfrac{3}{32}\xi^6(I_4+2I_6+I_8)+\ldots\},$$

and for $m \geqslant 3$

$$[5.11] \qquad \int_1^{u_0} \frac{du}{uf(u,\xi,m)} = 2\xi^{\frac{1}{2}}\left\{\theta+\frac{1}{2m}\xi^{m+1}(I_2+I_4+\ldots+I_{2m})+\ldots\right\}.$$

If the first two terms are retained in the expansions of [5.1] and [5.8] in powers of ξ^{m+1}, it follows that [3.11] simplifies to

$$[5.12] \quad \xi^{-\frac{1}{2}}\left[\tan\theta-\theta\left\{\frac{1}{4}+\frac{3}{4}\left(1+\frac{1}{m}\right)\xi\right\}\right]$$
$$+\frac{1}{2m}\xi^{m+\frac{1}{2}}\left[\frac{\theta}{2}+I_{2m}+\frac{3}{4}\left\{1-\left(1+\frac{1}{m}\right)\xi+\frac{1}{m}\xi^{m+1}\right\}(I_2+I_4+\ldots+I_{2m})\right]+\ldots = J(\psi_0,m),$$

where ξ is defined in [5.3], and we may also write

$$1-\left(1+\frac{1}{m}\right)\xi+\frac{1}{m}\xi^{m+1} = (1-\xi)^2\left[1+\frac{(m-1)}{m}\xi+\frac{(m-2)}{m}\xi^2+\ldots+\frac{1}{m}\xi^{m-1}\right].$$

The higher terms in the expansion, required for $m = 1$ and $m = 2$, are readily obtained from the relevant equations above. It is convenient to solve for θ rather than for ξ.

The cases $m \gg 1$ and $\xi \ll 1$ are of particular interest. The equation [5.12] becomes, very nearly,

$$[5.13] \qquad \xi^{-\frac{1}{2}}\left[\tan\theta-\theta\left\{\frac{1}{4}+\frac{3}{4}\left(1+\frac{1}{m}\right)\xi\right\}\right] = J(\psi_0,m)$$

and also [3.7] simplifies to

$$[5.14] \qquad Kh = 2\xi^{\frac{1}{2}}\theta = 2\xi^{\frac{1}{2}}\tan^{-1}[\xi\exp(-\beta z_1 e\psi_0)-1]^{\frac{1}{2}},$$

which can also be written in the form

$$[5.15] \qquad \psi_0-\psi_m = \frac{2kT}{z_1 e}\log\cos(\tfrac{1}{2}Kh\xi^{-\frac{1}{2}}).$$

This result was first given by Langmuir (17) and represents the limiting case where the potential ψ is very large at all positions in the diffuse layers of the two plates. The concentration of anions is so small compared with that of the cations that we may omit the term in n_2^0 in the Poisson–Boltzmann equation [2.1], which consequently has as solution equation [5.15]. This omission of the term in n_2^0 implies that either $m \gg 1$ or $\xi \ll 1$. If $-\psi_0$ is put equal to infinity, then it follows from [5.15] that

$$[5.16] \qquad \cos(\tfrac{1}{2}Kh\xi^{-\frac{1}{2}}) = 0 \quad \text{or} \quad \xi = (hK/\pi)^2$$

and so the pressure between the plates becomes

$$[5.17] \qquad p_e(h) = kTn_1^0\left(\frac{1}{\xi}-1\right) \approx \frac{\pi\epsilon}{8}\left(\frac{kT}{z_1e}\right)^2\frac{1}{h^2}$$

if $\xi \ll 1$. Since [5.14] can also be written as

$$[5.18] \qquad \xi\cos^2(\tfrac{1}{2}hK\xi^{-\frac{1}{2}}) = \exp(\beta z_1e\psi_0),$$

we see that to this approximation (applicable when h is small and ψ_0 is large), at a specified separation $2h$ the electrical force is expressed as a function of the surface potential ψ_0 and concentration n_1^0 of coagulation counter-ions only. This is the "zero-order" theory, which states that if ψ_0 is assumed independent of h, then the flocculating concentration depends only on the concentration of the ions oppositely charged to the plates and is independent of the concentration or nature of the ions carrying a charge of the same sign as the plates.

The equation [5.12] for θ, and therefore for ξ, has been solved numerically for $m = 1$, 2, 3, 4, and ∞. Substitution into [3.7] then yields Kh and hence, from [3.12], the quantity $(.1K/6\epsilon)(z_1e/kT)^2$ can be determined. This provides a measure of the concentration n_1^0 of the coagulating univalent ions. The results are given in Table II.

TABLE II

	$\psi_0 = -50$ mv					$\psi_0 = -100$ mv				
m	1	2	3	4	∞	1	2	3	4	∞
θ	0.885	0.872	0.865	0.864	0.880	1.23	1.23	1.23	1.24	1.26
ξ	0.338	0.327	0.322	0.321	0.334	0.168	0.166	0.168	0.170	0.190
Kh	1.08	1.01	0.98	0.98	1.02	1.02	1.01	1.01	1.02	1.09
$(KA/6\epsilon)\ (z_1e/kT)^2$	1.62	1.66	1.71	1.76	2.10	4.21	4.61	4.77	4.90	5.58
$(K_m/K_1)^2$	1	1.05	1.10	1.17	1.67	1	1.20	1.28	1.35	1.76
$2m/(1+m)$	1	1.33	1.50	1.60	2	1	1.33	1.50	1.60	2

According to the Derjaguin–Verwey–Overbeek theory, in which the Gouy–Chapman theory of the diffuse layer is applied for all distances from the particle wall and the surface potential ψ_0 is assumed independent of separation, the critical coagulating value of K and therefore of n_1^0 increases with the valency m of the non-flocculating ion. Denoting the value of K for the electrolyte type $1-m$ by K_m, the ratio $(K_m/K_1)^2$, which equals the corresponding ratio of the flocculating ion concentrations, is given in Table II. Ostwald's rule states that for a given sol $n_1^0(1+m)/m$ is a constant and therefore K is proportional to $\{m/(1+m)\}^{\frac{1}{2}}$; this yields $(K_m/K_1)^2 = 2m/(1+m)$, which is also given in the table. It is seen that Ostwald's rule is not verified by the Derjaguin–Verwey–Overbeek theory in the particular case of a $1-m$ binary electrolyte in which the flocculating ion is univalent, although the difference is not large.

It is possible to obtain a good approximation to the free energy $F(\infty)$ for a single plate when m is large ($\geqslant 4$ say) by employing the expansion [5.7]. Thus $F(\infty)$ is given by [3.9] where, from [3.10] and [5.7],

$$[5.19] \qquad J(\psi_0, m) = \tan\theta_0 - \theta_0 - \frac{1}{2m}\{\theta_0 - I_{2m}(\theta_0)\} + \ldots,$$

where $\sec^2\theta_0 = v_0 = \exp(-\beta z_1e\psi_0)$. The two terms on the right are calculated for $m = 3$ and 4 and $\psi_0 = -50$ and -100 mv in Table I. The first term on the right is identical with [4.8] and gives the exact result for $m = \infty$.

6. APPLICATION OF STERN MODEL IN THE ABSENCE OF SPECIFIC ADSORPTION

Only one of the two modifications which Stern introduced in the Gouy–Chapman theory will be considered here, namely, the effect of the finite ion size, which is sufficient when specific adsorption of ions on the plate surface is absent. It is assumed that there is a layer of water molecules at the wall which is free of ions, the thickness d, say, of this layer being the minimum distance of approach of the ions of the coagulating electrolyte. Experimental evidence (19, 20, 21) indicates that the water in this inner region of thickness d is partially oriented even in the absence of any charge on the surface and as a result of this partial dielectric saturation, the effective dielectric constant of this inner region, which will be denoted by ϵ_1, is found to be much less than that of bulk water. If $\Psi = \Psi(\sigma)$ is the potential at the colloidal wall, $x = h$ when the plate charge density is σ, and $\Psi_d = \Psi_d(\sigma)$ that at the other boundary, $x = h-d$ of the inner region (Fig. 2) then we may assume the relation

$$\Psi - \Psi_d = \frac{4\pi d\sigma}{\epsilon_1} + \chi_d, \qquad [6.1]$$

where χ_d is the potential drop due to the dipole orientation in the absence of surface charge σ; this contributes to the so-called Lange χ-potential. It will be shown in a later paper that ϵ_1 should be interpreted as a differential rather than an integral dielectric constant. The limit of the diffuse layer is now assumed to be situated at $x = h-d$ instead of $x = h$, as in the Gouy–Chapman theory.

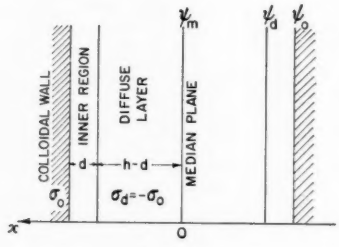

FIG. 2.

To determine the free energy for the above model of the electric double layers, it is necessary to discuss the origin of the first equation in [2.5]. The free energy consists of two terms, an 'electrical' part $F_e(h)$ given by

$$F_e(h) = 2\int_0^{\sigma_0} \Psi d\sigma \qquad [6.2]$$

and a chemical part F_c, which takes the form

$$F_c = 2C\sigma_0, \qquad [6.3]$$

where C is a constant, when the potential-determining ions responsible for the surface charge form part of the particle lattice. Since $F(h)$ must be a minimum with respect to transfer of potential-determining ions from the electrolyte dispersion medium to the plates,

$$\frac{dF(h)}{d\sigma_0} = \frac{d}{d\sigma_0}\Big[F_e(h) + F_c\Big] = 0, \qquad [6.4]$$

or, since $\psi_0 = \Psi(\sigma_0)$,

$$\psi_0 + C = 0. \tag{6.5}$$

Integrating by parts

$$F(h) = F_e(h) + F_c = 2(\psi_0 + C)\sigma_0 - 2\int_0^{\psi_0} \sigma d\Psi = -2\int_0^{\psi_0} \sigma d\Psi. \tag{6.6}$$

Substituting [6.1] we obtain

$$F(h) = -2\int_0^{\psi_d} \sigma d\Psi_d - \frac{4\pi d\sigma_0^2}{\epsilon_1}, \tag{6.7}$$

where $\psi_d = \Psi_d(\sigma_0)$ is the potential at the limit of the diffuse layer $x = h - d$ when the plate is fully charged. The last result was given by Mackor (22). Levine (23) considered the more general case where the inner region contains an adsorbed layer of counter-ions and his formula reduces to [6.6] in the case of no specific adsorption.

The charge density σ is equal in magnitude and opposite in sign to the charge σ_d, say, contained in a column of the diffuse layer of unit cross section and situated between the planes $x = 0$ and $x = h - d$. It follows that, apart from the constant factor d/ϵ_1, the energy expression [6.7] depends solely on the charge distribution in the diffuse layer, and this is also true of the formula in [2.5]. Furthermore, the double layer repulsion per unit area can be written as

$$p_e(h') = -\frac{1}{2}\left[\frac{\partial F_e(h)}{\partial h}\right]_{\sigma_0} = \left[-\frac{\partial}{\partial h}\int_0^{\sigma_0} \Psi_d d\sigma\right]_{\sigma_0}, \tag{6.8}$$

where $h' = h - d$, and [6.1] is substituted into [6.2]. The proof of the first relation in [6.8] has been given by Verwey and Overbeek (2), Levine (24), and Bell and Levine (25). Since $\sigma = -\sigma_d$, this means that the force per unit area also depends on the charge distribution in the diffuse layer only and thus is given by the last expression in [2.6]. Indeed it has already been stressed (26, 27) that the latter force formula, which is expressed in terms of the ion densities at the median plane, is always valid, being independent of the particular conditions at the particle surface, *provided* the median plane is embedded in the diffuse layers and the Gouy–Chapman equations can be applied to the diffuse layer. It follows that the first term on the right of [6.7] may be equated to the expression [2.9] if h is replaced by $h' = h - d$ and ψ_0 by ψ_d, and so

$$F(h) = -2E_e(h') - 2h'p_e(h') - \frac{4\pi d\sigma_0^2}{\epsilon_1}. \tag{6.9}$$

The van der Waals attraction will take a different form from [2.10] since the inner region of water not only has a different dielectric constant but is probably in a state of considerable electrostriction. It is convenient to write the van der Waals energy in the form

$$W(h') = -\frac{A}{48\pi h'^2} + A'F_w(h'), \tag{6.10}$$

where the second term is the correction due to the difference between the materials of the inner region and the colloidal plates. Certainly for large separations the first term on the right will be the dominant one and in view of the mathematical complexities

it seems sufficient for the present purposes to neglect the second term. In this case, the two stability conditions [2.12] become

$$[6.11] \qquad p_e(h') = \frac{A}{48\pi h'^3}, \qquad 2E_e(h')+3h'p_e(h')+F'(\infty)+\frac{4\pi d\sigma_0^2}{\epsilon_1} = 0,$$

where $F'(\infty)$ is the free energy per unit area of a single plate.

It is convenient to introduce

$$[6.12] \qquad v_d = \exp(-\beta z_1 e\psi_d), \qquad u_d = \xi v_d, \qquad v_d(\infty) = \exp\{-\beta z_1 e\psi_d(\infty)\},$$

where ξ is already defined in [3.2], and noting that ψ_d varies with separation $2h$, its value at infinite separation is denoted by $\psi_d(\infty)$. Then $E_e(h')$ is given by [3.5] and similarly Kh' by [3.7], provided the upper limit of integration u_0 on the right-hand side of both equations is replaced by u_d. Also $p_e(h')$ is again given by [3.6]. If σ_0 has the value $\sigma_0(\infty)$ at infinite separation and $J(\psi_d(\infty), m)$ is defined by [3.10], when the upper limit of integration is replaced by $v_d(\infty)$, then

$$[6.13] \qquad F'(\infty) = -8n_1^0 kTJ(\psi_d(\infty), m)-\frac{4\pi d\sigma_0^2(\infty)}{\epsilon_1}.$$

By [3.4] the charge density σ_0 can be written as

$$[6.14] \qquad \sigma_0 = -\left(\frac{n_1^0 kT\epsilon}{2\pi}\right)^{\frac{1}{2}} f(u_d, \xi, m)$$

and the corresponding form for $\sigma_0(\infty)$ is obtained by replacing u_d by $v_d(\infty)$ and ξ by 1. Thus the second condition in [6.11] reads

$$[6.15] \qquad \frac{1}{2}\int_1^{u_d}\frac{f(u, \xi, m)}{u}du+\frac{3}{8}\left(\frac{1}{m}\xi^m+\xi^{-1}-\frac{(m+1)}{m}\right)\int_1^{u_d}\frac{du}{uf(u, \xi, m)}-J(\psi_d(\infty), m)$$
$$= \frac{Kd\epsilon}{4\epsilon_1}[f^2(v_d(\infty), 1, m)-f^2(u_d, \xi, m)]$$

after dividing through by $8n_1^0 kT/K$. This relation becomes identical with [3.11] when the thickness d of the inner region tends to zero. For non-vanishing d, however, it differs from [3.11] in that the electrolyte concentration variable has not been eliminated and it is no longer possible to solve for ξ in terms of u_d without introducing the other stability condition in [6.11].

It is convenient to express the condition [6.1] as

$$[6.16] \qquad -\beta z_1 e\psi_d = \log v_d = \Phi_0-\frac{Kd\epsilon}{\epsilon_1}f(u_d, \xi, m),$$

where $\Phi_0 = \beta z_1 e(\chi_d-\psi_0)$, which is constant independent of d. If use is made of [3.3], then at $h = \infty$ (when $\xi = 1$) this relation can be put in the form

$$[6.17] \qquad \log v_d(\infty)+\frac{Kd\epsilon}{\epsilon_1}[v_d^{\frac{1}{2}}(\infty)-v_d^{-\frac{1}{2}}(\infty)]H^{\frac{1}{2}}(v_d(\infty), m) = \Phi_0$$

where

$$[6.18] \qquad H(s, m) = 1+\frac{(m-1)}{m}\cdot\frac{1}{s}+\frac{(m-2)}{m}\cdot\frac{1}{s^2}+\ldots+\frac{1}{ms^{m-1}}.$$

Thus

$$H(s,1) = 1, \qquad H(s,2) = 1+\frac{1}{2s}, \qquad H(s,3) = 1+\frac{2}{3s}+\frac{1}{3s^2},$$

$$H(s,4) = 1+\frac{3}{4s}+\frac{1}{2s^2}+\frac{1}{4s^3}, \qquad H(s,\infty) = \frac{s}{(s-1)}.$$

The equation [6.17] yields $v_d(\infty)$ (or $\psi_d(\infty)$) as a function of K. Now χ_d measures a residual potential drop in the absence of surface charge and will be assumed independent of electrolyte type (i.e. of m). Since ψ_0 also does not depend on valency type we may presume Φ_0 to be independent of m. The thickness d of the Stern layer is some mean of the two distances of nearest approach to the colloidal wall of the cation and anion, which we denote by d_I and d_{II} respectively. For simplicity, let $d_I = d_{II}$ for the 1–1 electrolyte and noting that the radius of the hydrated anion will in general increase with its hydration number let d_{II} increase with its valency m. Then it is likely that for small $m > 1$, $d > d_I$ but the following considerations suggest that as m increases d diminishes again, tending to d_I. If the hydration number is assumed proportional to m then the ratio d_{II}/d_I will be roughly proportional to $m^{1/3}$ whereas the ratio of the concentration of anions to cations at a distance $x = d_{II}$, where the potential is $\psi_{II}(<0)$ say, will be $(1/m)\exp[\beta e(m+1)\psi_{II}]$, which rapidly tends to zero with increasing m. This means that the number of anions that would be present in the region $d_I < x < d_{II}$ if they were not excluded through steric hindrance, would be so small for $m \geqslant 3$, say, that the distance d is determined by the cations only and is therefore equal to d_I. This behavior of d as a function of m will be borne in mind below.

To examine the relation [6.16] for general h we introduce $u_d = \sec^2\alpha$ and therefore $v_d = \sec^2\alpha/\xi$. Then, making use of [3.3], [6.16] can be written as

$$[6.19] \quad \log\left(\frac{\sec^2\alpha}{\xi}\right)+\frac{Kd\epsilon}{\epsilon_1}\xi^{-\frac{1}{2}}\tan\alpha\left[1-\frac{1}{m}\xi^{m+1}\cos^{2m}\alpha\{1+\sec^2\alpha+\ldots+\sec^{2(m-1)}\alpha\}\right]^{\frac{1}{2}} = \Phi_0.$$

Similarly, equation [6.15] reads:

$$[6.20] \quad \xi^{-\frac{1}{2}}\left[\tan\alpha-\alpha\left\{\frac{1}{4}+\frac{3}{4}\left(1+\frac{1}{m}\right)\xi\right\}\right]$$
$$+\frac{1}{2m}\xi^{m+\frac{1}{2}}\left[\frac{\alpha}{2}+I_{2m}+\frac{3}{4}\left\{1-\left(1+\frac{1}{m}\right)\xi+\frac{1}{m}\xi^{m+1}\right\}(I_2+I_4+\ldots+I_{2m})\right]$$
$$+\ldots-J(\psi_d(\infty),m) = \frac{Kd\epsilon}{4\epsilon_1}\left[\frac{(v_d(\infty)-1)^2}{v_d(\infty)}H(v_d(\infty),m)\right.$$
$$\left.-\xi^{-1}\tan^2\alpha\left\{1-\frac{1}{m}\xi^{m+1}\cos^{2m}\alpha(1+\sec^2\alpha+\ldots+\sec^{2(m-1)}\alpha)\right\}\right],$$

where, as in [5.12] only two terms in the expansion in powers of ξ^{m+1} have been retained on the left-hand side. In the three equations [6.17], [6.19], and [6.20], there are four unknowns α, ξ, K, and $\psi_d(\infty)$ (or $v_d(\infty)$) so that one more equation is required; this is provided by the first stability condition in [6.11]. Making use of [3.7] and [5.11], this condition is equivalent to [3.12] provided h, u_0, and θ are replaced by h', u_d, and α. The required fourth relation is therefore

$$[6.21] \quad M = C\xi^{\frac{1}{2}}\left\{1-\left(1+\frac{1}{m}\right)\xi+\frac{1}{m}\xi^{m+1}\right\}\left[\alpha+\frac{1}{2m}\xi^{2m+1}(I_2+I_4+\ldots+I_{2m})+\ldots\right]^3,$$

where

$$[6.22] \qquad M = \frac{Kd\epsilon}{\epsilon_1}, \qquad C = \frac{48d\epsilon^2}{A\epsilon_1}\left(\frac{kT}{z_1e}\right)^2,$$

and the I_n are functions of α in [6.20] and [6.21].

The simplest case is $m = \infty$. Attaching a suffix ∞ to the quantities M and C and introducing $\phi = -\beta z_1 e\psi_d(\infty) > 0$ the four equations are readily obtained, namely

$$M_\infty = \frac{\Phi_0-\phi}{(e^\phi-1)^{\frac{1}{2}}} = \frac{\Phi_0-\log(\sec^2\alpha/\xi)}{\xi^{-\frac{1}{2}}\tan\alpha} = C_\infty\xi^{\frac{1}{2}}\alpha^3(1-\xi),$$

$$[6.23] \qquad \xi^{-\frac{1}{2}}\left[\tan\alpha-\frac{\alpha}{4}(1+3\xi)\right]-J(\psi_d(\infty),\infty) = \frac{M_\infty}{4}\left[e^\phi-1-\xi^{-1}\tan^2\alpha\right].$$

For $m = 1$, attaching the suffix 1 to the quantities M and C, the corresponding equations are

$$[6.24] \qquad M_1 = \frac{\Phi_0-\phi}{2\sinh\phi/2} = \frac{\Phi_0-\log(\sec^2\alpha/\xi)}{\xi^{-\frac{1}{2}}\tan\alpha(1-\xi^3\cos^2\alpha)^{\frac{1}{2}}}$$

$$= C_1\xi^{\frac{1}{2}}(1-\xi)^2\left[\alpha+\frac{1}{2}\xi^2I_2+\frac{3}{8}\xi^4I_4+\ldots\right]^3$$

and

$$[6.25] \qquad \xi^{-\frac{1}{2}}\left[\tan\alpha-\frac{\alpha}{4}(1+6\xi)\right]+\frac{1}{2}\xi^{3/2}\left[\frac{\alpha}{2}+I_2+\frac{3}{4}(1-\xi)^2I_2\right]$$

$$+\frac{1}{8}\xi^{7/2}\left[I_4-I_2+\frac{9}{4}(1-\xi)^2I_4\right]+\ldots-J(\psi_d(\infty),1)$$

$$= \frac{M_1}{4}\left[4\sinh^2\frac{\phi}{2}-\xi^{-1}\tan^2\alpha(1-\xi^2\cos^2\alpha)\right],$$

where three terms in the expansion in the powers of $\xi^{m+1} = \xi^{3/2}$ on the left-hand side of [6.20] have been retained. For $m = 2$ we obtain

$$[6.26] \qquad M_2 = \frac{\Phi_0-\phi}{2\,(1+\frac{1}{2}e^{-\phi})^{\frac{1}{2}}\sinh\phi/2} = \frac{\Phi_0-\log(\sec^2\alpha/\xi)}{\xi^{-\frac{1}{2}}\tan\alpha[1-\frac{1}{2}\xi^3(\cos^2\alpha+\cos^4\alpha)]^{\frac{1}{2}}}$$

$$= C_2\xi^{\frac{1}{2}}(1-\xi)^2\left(1+\frac{1}{2}\xi\right)\left[\alpha+\frac{1}{4}\xi^3(I_2+I_4)+\ldots\right]^3$$

and

$$[6.27] \qquad \xi^{-\frac{1}{2}}\left[\tan\alpha-\frac{\alpha}{8}(2+9\xi)\right]+\frac{1}{4}\xi^{5/2}\left[\frac{\alpha}{2}+I_4+\frac{3}{8}(1-\xi)^2(2+\xi)(I_2+I_4)\right]+\ldots$$

$$-J(\psi_d(\infty),2) = \frac{M_2}{4}\left[4\sinh^2\frac{\phi}{2}\,(1+\tfrac{1}{2}e^{-\phi})-\xi^{-1}\tan^2\alpha\{1-\tfrac{1}{2}\xi^3(\cos^2\alpha+\cos^4\alpha)\}\right].$$

Some numerical solutions of the above equations are given in Table III for $\Phi_0 = 10$ ($|\psi_0-\chi_d| = 250$ mv). An examination of [6.17] shows that for specified Φ_0 and Kd the magnitude of the potential $\psi_d(\infty)$ diminishes with increasing m. This suggests that if

the value $\phi = 2$ ($\psi_d(\infty) = -50$ mv), for example, is chosen for $m = 1$ then $\phi < 2$ for $m > 1$. The subscripts 1, 2, or ∞ have been assigned to various quantities to designate that $m = 1$, 2, or ∞ respectively. Values of ϕ have been chosen to provide a range of values for the ratio d_m/d_1, of the thicknesses of the inner region in 1–1 and 1–m electrolytes. From the definition of C in [6.22], $d_m/d_1 = C_m/C_1$ and the corresponding ratio of the coagulating ion concentrations is $(K_m/K_1)^2 = (M_m d_1/M_1 d_m)^2$. It is seen that

TABLE III

m		1	2			∞	
ϕ		2	2	1.9	1.96	1.8	1.75
M_m		3.40	3.29	3.55	3.40	3.65	3.78
α		0.917	0.920	0.870		0.864	0.857
ξ		0.310	0.305	0.285		0.293	0.302
C_m		14.9	13.4	16.9	14.9	14.8	15.7
d_m/d_1			0.90	1.14	1	0.99	1.06
$\left(\frac{K_m}{K_1}\right)^2$	Stern		1.16	0.84	1.00	1.17	1.11
	Gouy–Chapman		1.05			1.67	($\phi = 2$)
	Ostwald (any ϕ)			1.33		2	

if $d_m \geqslant d_1$ then the Stern theory yields a greater divergence from Ostwald's rule than the Gouy–Chapman theory. In view of the variation with sol concentration of the coagulating ion activity coefficient, as stressed by Weiser and Tezak, at best the Ostwald rule seems to be an empirical approximation. On the other hand, it is probable that for asymmetrical electrolytes where $m \geqslant 3$, the Poisson–Boltzmann equation, which is the basis of the Gouy–Chapman–Stern theory of the diffuse layer, is also inadequate because of the neglect, for example, of the fluctuation or self-atmosphere effect. It is well known that the limitations of the classical Debye–Huckel theory for electrolytes containing ions of high valency are largely attributed to this fluctuation term. Since the theory developed in this paper applies to very dilute sols whereas Ostwald's rule apparently holds for concentrated sols, the measure of agreement between the Gouy–Chapman–Stern theory and experiment for the valency type 1–m is not easily determined.

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THE SULPHURIC ACID SOLVENT SYSTEM

PART I. ACID–BASE REACTIONS[1]

R. H. FLOWERS, R. J. GILLESPIE, AND E. A. ROBINSON

ABSTRACT

Acid–base reactions in the solvent sulphuric acid are discussed. Such reactions are conveniently studied by electrical conductivity measurements. A relation between the composition at which the conductivity has a minimum value and the strengths of the acid and base is derived. Values of the dissociation constants of acids and bases obtained in this way are shown to be in good agreement with values obtained by other methods.

INTRODUCTION

Sulphuric acid is an amphoteric solvent that undergoes fairly extensive autoprotolysis (1, 2).

$$2H_2SO_4 = H_3SO_4^+ + HSO_4^-$$

An acid of the sulphuric acid system may be defined as any substance that gives rise to the sulphuric acidium ion $H_3SO_4^+$. A simple monoprotic acid would ionize according to the equation

$$HA + H_2SO_4 = H_3SO_4^+ + A^-.$$

A base may be defined as any substance that gives rise to the hydrogen sulphate ion HSO_4^-. A simple monoprotic base would ionize according to the equation

$$B + H_2SO_4 = BH^+ + HSO_4^-.$$

Reaction between an acid and a base occurs according to the equation

$$H_3SO_4^+ + HSO_4^- = 2H_2SO_4,$$

which is, of course, just the reverse of the autoprotolysis reaction.

Because of the abnormally high mobilities of the $H_3SO_4^+$ and HSO_4^- ions the electrical conductivity of any solutions can be regarded, to a reasonably good approximation, as being due to these ions alone (1, 2). In acid–base reactions in which, for example, a base is added to a solution of an acid, the concentration of $H_3SO_4^+$ decreases and the concentration of HSO_4^- increases, causing the conductivity to decrease initially, pass through a minimum, and increase again. The position of minimum conductivity depends on the strengths of the acid and base and on their modes of ionization. By following the change in conductivity accompanying an acid–base reaction, information on the nature of the reaction and on the strength of the acid and base may be obtained. This paper gives a general treatment of the variation of the electrical conductivity during an acid–base reaction and some examples of the determination of acid and base strengths by this method are discussed.

THEORY

The detailed treatment of acid–base reactions is somewhat complicated by the existence of another form of self-dissociation of the solvent in addition to autoprotolysis, namely self-dehydration (1, 2), which may be summarized by the equations

$$2H_2SO_4 = H_2S_2O_7 + H_2O,$$
$$H_2S_2O_7 + H_2SO_4 = H_3SO_4^+ + HS_2O_7^-,$$
$$H_2O + H_2SO_4 = H_3O^+ + HSO_4^-.$$

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Contribution from the Department of Chemistry, University College, London, W.C.1, England, and the Department of Chemistry, McMaster University, Hamilton, Ontario.

The concentrations of all the species in a solution containing an acid HA and a base B are related by the following expressions: the four equilibrium constants describing the self-dissociation of the solvent, whose values are given in Table I (3, 4),

$$K_{ap} = [H_3SO_4^+][HSO_4^-]^* \quad [1]$$

$$K_{id} = [H_3O^+][HS_2O_7^-] \quad [2]$$

$$K_{H_2S_2O_7} = [H_3SO_4^+][HS_2O_7^-]/[H_2S_2O_7] \quad [3]$$

$$K_{H_2O} = [H_3O^+][HSO_4^-]/[H_2O], \quad [4]$$

the dissociation constants of the acid and the base,

$$K_a = [H_3SO_4^+][A^-]/[HA] \quad [5]$$

$$K_b = [BH^+][HSO_4^-]/[B], \quad [6]$$

the condition for electrical neutrality,

$$[BH^+] + [H_3O^+] + [H_3SO_4^+] = [A^-] + [HS_2O_7^-] + [HSO_4^-], \quad [7]$$

and the condition that the composition of the solvent is H_2SO_4,

$$[H_3O^+] + [H_2O] = [HS_2O_7^-] + [H_2S_2O_7]. \quad [8]$$

TABLE I

Equilibrium constants for the self-dissociation reactions of sulphuric acid

	10° C	25° C
$K_{ap} = [H_3SO_4^+][HSO_4^-]$	1.7×10^{-4}	2.4×10^{-4}
$K_{id} = [H_3O^+][H_2S_2O_7^-]$	3.0×10^{-5}	4.2×10^{-5}
$K_{H_2S_2O_7} = [H_3SO_4^+][HS_2O_7^-]/[H_2S_2O_7]$	1.4×10^{-2}	1.4×10^{-2}
$K_{H_2O} = [H_3O^+][HSO_4^-]/[H_2O]$	1	1

For the special case where $H_2S_2O_7$ is the acid equation [5] is not needed and equations [7] and [8] are replaced by

$$[BH^+] + [H_3O^+] + [H_3SO_4^+] = [HS_2O_7^-] + [HSO_4^-] \quad [9]$$

and

$$m^s_{H_2S_2O_7} = [HS_2O_7^-] + [H_2S_2O_7] - [H_3O^+] - [H_2O] \quad [10]$$

where $m^s_{H_2S_2O_7}$ is the stoichiometric molal concentration of $H_2S_2O_7$.

For the special case where water is the base being neutralized equation [6] is not needed and [7] and [8] are replaced by

$$[H_3O^+] + [H_3SO_4^+] = [A^-] + [HS_2O_7^-] + [HSO_4^-] \quad [11]$$

and

$$m^s_{H_2O} = [H_3O^+] + [H_2O] - [H_2S_2O_7] - [HS_2O_7^-] \quad [12]$$

where $m^s_{H_2O}$ is the stoichiometric molal concentration of water. The concentrations of all the species for any concentrations of acid and base in an acid–base reaction may be calculated from these equations by the method that has been previously described (3).

The acids of the sulphuric acid system, e.g. tetra(hydrogensulphato) boric acid, $HB(HSO_4)_4$ (5), and disulphuric acid, $H_2S_2O_7$, are either known only in solution in sulphuric acid, or are only easily handled in solution, whereas the common bases, e.g.

**All concentrations in this paper are in molal units and* [A] *and* m_A *are both used to denote the molal concentration of the species A.*

potassium sulphate, benzoic acid, and nitrobenzene, are easily obtained in the pure state and they are easily handled as such. Thus it is convenient to carry out acid–base reactions by adding the pure base to a solution of the acid. The position of minimum conductivity is then most conveniently expressed in terms of the ratio $r = n_b/n^1_a$ of the number of moles of added base, n_b, to the initial number of moles of acid n^1_a. Figures 1 and 2 show

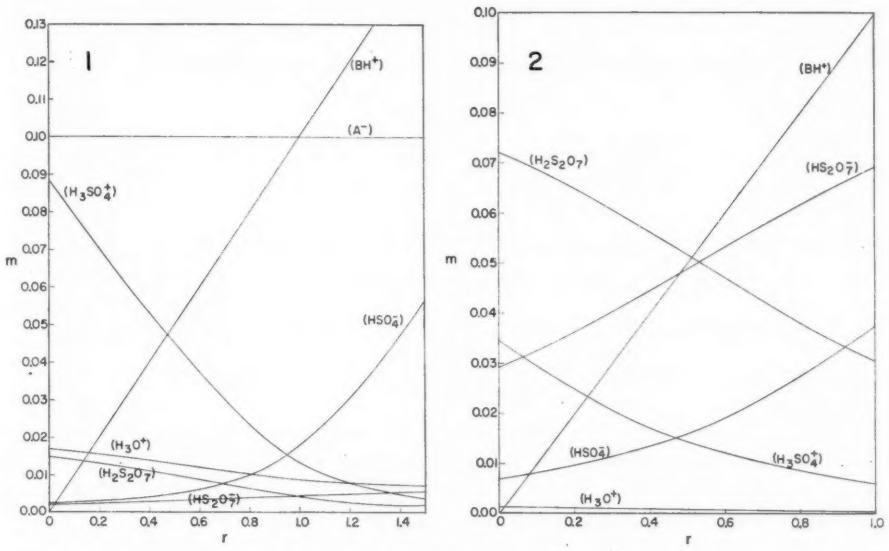

FIG. 1. Concentrations of species in a strong acid – strong base reaction at 25° ($m^1_{HA} = 0.1$).
FIG. 2. Concentrations of species in a disulphuric acid – strong base reaction at 25° ($m^1_{H_2S_2O_7} = 0.1$).

how the concentrations of all the species in reactions of the acids HA and $H_2S_2O_7$ with a strong base B vary with r for an initial concentration of acid $m^1_{H_2S_2O_7}$ or $m^1_{HA} = 0.1$.

The conductivity of the solution at any point in such a reaction is given by

$$\kappa = 10^{-3}\rho([H_3SO_4^+]\lambda_{H_3SO_4^+} + [HSO_4^-]\lambda_{HSO_4^-} + [H_3O^+]\lambda_{H_3O^+} + [HS_2O_7^-]\lambda_{HS_2O_7^-} + [BH^+]\lambda_{BH^+} + [A^-]\lambda_{A^-}) \quad [13]$$

where the λ's represent ion mobilities (molar conductances) and where it has been assumed that molar concentrations c may be replaced by $m\rho$ where m is the molal concentration and ρ the density of the solution. In view of the very small conductivities of all normally conducting ions (2) it is a reasonably good approximation to neglect their contribution to the conductivity of the solution in comparison with the relatively large contribution made by the highly conducting $H_3SO_4^+$ and HSO_4^- ions; hence we may write

$$\kappa = 10^{-3}\rho\,(\lambda_+ m_+ + \lambda_- m_-) \quad [14]$$

where λ_+ and λ_- are the mobilities, and m_+ and m_- are the molal concentrations of $H_3SO_4^+$ and HSO_4^- respectively. The mobilities of $H_3SO_4^+$ and HSO_4^- decrease with increasing concentration but at total ion concentrations of the order of $0.2m$, such as occur in the experiments described later, the variation of these mobilities with concentration is small (6) and is unlikely to be sufficiently large to appreciably affect the position of the

conductivity minimum. The ratio $\mu = \lambda_+/\lambda_-$ is, at least approximately, independent of concentration and has the value 1.45 (4). Hence we may write

$$\kappa = 10^{-3}\,\lambda_-(m_+\mu + K_{ap}/m_+). \qquad [15]$$

Since μ is constant and assuming that λ_- is also constant we have

$$\frac{\partial \kappa}{\partial m_+} = 10^{-3}\rho\lambda_-(\mu - K_{ap}/m_+^{\,2}) \qquad [16]$$

At the composition of minimum conductivity $\partial\kappa/\partial m_+ = 0$ and therefore

$$\mu - K_{ap}/(m_+)^2_{min} = 0,$$

or

$$(m_+)_{min} = \sqrt{K_{ap}/\mu}, \qquad (m_-)_{min} = \sqrt{K_{ap}\mu}. \qquad [17]$$

Substituting the values of K_{ap} given in Table I and $\mu = 1.45$ we find

$$(m_+)_{min} = 0.0132 \text{ at } 25° \quad \text{and} \quad 0.0108 \text{ at } 10°$$

$$(m_-)_{min} = 0.0182 \text{ at } 25° \quad \text{and} \quad 0.0157 \text{ at } 10°$$

Base–HA *Reactions*

Using the relation

$$[HS_2O_7^-] = \left(\frac{K_{id}\cdot K_{H_2S_2O_7}}{[H_3SO_4^+] + K_{H_2S_2O_7}}\right)^{1/2} \qquad [18]$$

which can be derived from equations [1] to [3] we obtain

$$[HS_2O_7^-]_{min} = 0.0045 \text{ at } 25° \text{ and } 0.0041 \text{ at } 10°$$

and from equation [2]

$$[H_3O^+]_{min} = 0.0088 \text{ at } 25° \text{ and } 0.0075 \text{ at } 10°.$$

Since

$$[BH^+] = m_b/(1 + [HSO_4^-]/K_b)$$

where m_b is the molality of added base B

and

$$[A^-] = m^1_a/(1 + [H_3SO_4^+]/K_a)$$

where m^1_a is the initial molality of the acid HA and if we write

$$y = [BH^+] - [A^-] = [HSO_4^-] + [HS_2O_7^-] - [H_3SO_4^+] - [H_3O^+]$$

we have

$$r_{min} = (m_b)_{min}/m^1_a = (1 + [HSO_4^-]_{min}/K_b)\{(1 + [H_3SO_4^+]_{min}/K_a)^{-1} + y/a\} \qquad [19]$$

where $a = m^1_a$.

Substituting the values given above for 25° we have

$$r_{min} = (1 + 0.018/K_b)\{(1 + 0.013/K_a)^{-1} + 0.0007/a\}. \qquad [20]$$

Values of r_{min} for various values of K_a and K_b are given in Table II for $a = 0.1m$.

TABLE II

Calculated values of r_{min} ("end points") for acid–base reactions

Strong acid ($K_a = \infty$)					
K_b	∞	1.0	10^{-1}	10^{-2}	10^{-3}
r_{min}	1.01	1.02	1.19	2.82	19.1
Strong base ($K_b = \infty$)					
K_a	∞	1.0	10^{-1}	10^{-2}	10^{-3}
r_{min}	1.01	0.98	0.89	0.44	0.08
Disulphuric acid					
K_b	∞	1.0	10^{-1}	10^{-2}	10^{-3}
r_{min}	0.56	0.57	0.66	1.58	11.3
Water					
K_a	∞	1.0	0.3	10^{-1}	10^{-2}
r_{min}	1.00	1.00	1.00	0.93	0.42

Base–$H_2S_2O_7$ Reactions

For reactions of $H_2S_2O_7$, $[HS_2O_7^-]_{min}$ can be calculated from the expression

$$[HS_2O_7^-] = \frac{a + \{a^2 + 4K_{id}(1 + [H_3SO_4^+]/K_{H_2S_2O_7})\}^{1/2}}{2(1 + [H_3SO_4^+]/K_{H_2S_2O_7})} \qquad [21]$$

where $a = m^1_{H_2S_2O_7}$ and which can be derived from equations [2], [3], and [10]. Then $[H_3O^+]_{min}$ can be obtained from equation [2] and $[BH^+]_{min}$ from equation [9]. For $a = 0.1$, for example, we thus obtain

$$[H_3O^+]_{min} = 0.0008, \quad [HS_2O_7^-]_{min} = 0.0519, \quad \text{and } [BH^+]_{min} = 0.0561.$$

Now we have

$$r_{min} = \frac{[BH^+]_{min}}{a}\{1 + [HSO_4^-]_{min}/K_b\}. \qquad [22]$$

Hence for $a = 0.1$ at 25°

$$r_{min} = 0.56\ (1 + 0.018/K_b). \qquad [23]$$

Table II gives values of r_{min} for various illustrative values of K_b.

H_2O–*Acid Reactions*

For reactions of H_2O and an acid HA, $[H_3O^+]_{min}$ can be calculated from the expression

$$[H_3O^+]_{min} = \{([A^-] + 0.0050) + \{([A^-] + 0.0050)^2 + 4K_{id}\}^{\frac{1}{2}}\}/2, \qquad [24]$$

which can be derived from equations [2] and [11] and the values of $[H_3SO_4^+]_{min}$ and $[HSO_4^-]_{min}$ given above, and where

$$[A^-] = a \Big/ \left(\frac{[H_3SO_4^+]_{min}}{K_a} + 1\right). \qquad [25]$$

Then $[HS_2O_7^-]_{min}$ can be obtained from equation [2], $[H_2S_2O_7]$ from equation [3], and $m^s_{H_2O}$ from equation [12]. Thus, for example, for $HB(HSO_4)_4$ ($K_a = 0.3$) and $a = 0.1$ we have $[H_3O^+]_{min} = 0.1012$ and $[HSO_7^-]_{min} = [H_2S_2O_7]_{min} = 0.0004$. Hence $(m^s_{H_2O})_{min} = 0.1004$ and $r_{min} = (m^s_{H_2O})_{min}/a = 1.004$. Values of r_{min} for various values of K_a are given in Table II.

The conductivity minimum in the special case of the H_2O–$H_2S_2O_7$ reaction has been discussed in detail in an earlier paper (4).

RESULTS

$H_2S_2O_7$ – *Strong Base Reactions*

The variation in the conductivity during the reactions of the strong bases $KHSO_4$ and benzoic acid with disulphuric acid solutions of various concentrations is shown in Fig. 3.

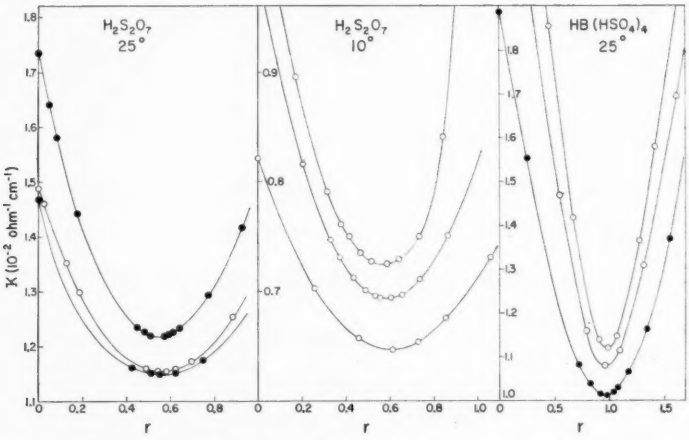

FIG. 3. Conductimetric titrations of strong bases with the acids $H_2S_2O_7$ and $HB(HSO_4)_4$. ○ $KHSO_4$, ● benzoic acid.

These experiments may be regarded as conductimetric acid–base "titrations" although the purpose of these titrations is not, as is usual, to estimate the concentration of either the solution of the acid or the base but rather to determine the dissociation constant of the acid or the base. In the present case the strengths of both the acid and the base are known so we may test the theory given above by comparing the predicted value of the "end point", i.e. the position of minimum conductivity, r_{min}, calculated from equation [23] with the observed value as is done in Table III. The agreement between observed and calculated values is excellent at 25° and good at 10°. The somewhat poorer agreement at 10° is probably due to the relatively greater contribution of the normally conducting ions to the conductivity at this temperature. This good agreement between the calculated and observed positions of the end point, r_{min}, provides support for the approximations made in the above theory and enables the equations to be used with confidence for the estimation of the strengths of weak acids and bases.

$HB(HSO_4)_4$ – *Strong Base Reactions*

The variation in the conductivity during the reactions of the strong bases $KHSO_4$ and benzoic acid with solutions of $HB(HSO_4)_4$ of various concentrations is shown in Fig. 3. The observed "end points", i.e. values of r_{min}, are given in Table III together with values for K_a calculated by means of equation [20]. The mean value, $K_a = 0.3$, is in good agreement with the value obtained from the conductivities of $HB(HSO_4)_4$ solutions (7) and

TABLE III

Disulphuric acid – strong base reactions					$HB(HSO_4)_4$ – weak base reactions (25° C)			
Base	a	r_{min}^{obs}	r_{min}^{calc}	T	Base	a	r_{min}	K_b
Benzoic acid	0.095	0.56	0.56	25°	*p*-Nitrotoluene	0.107	1.15	0.095
$KHSO_4$	0.099	0.56	0.56	25°	*o*-Nitrotoluene	0.050	1.22	0.071
Benzoic acid	0.154	0.55	0.55	25°	*m*-Nitrotoluene	0.118	2.00	0.017
$KHSO_4$	0.060	0.61	0.63	10°	*m*-Nitrotoluene	0.225	2.00	0.017
$KHSO_4$	0.112	0.59	0.61	10°	Nitrobenzene	0.116	2.90	0.009
$KHSO_4$	0.148	0.58	0.60	10°				

$HB(HSO_4)_4$ – strong base reactions (25° C)				$H_2S_2O_7$ – weak base reactions (25° C)			
Base	a	r_{min}	K_a	Base	a	r_{min}	K_a
Benzoic acid	0.079	0.97	0.3	*m*-Nitrotoluene	0.102	0.94	0.023
$KHSO_4$	0.114	0.97	0.3		0.153	0.94	0.023
$KHSO_4$	0.143	0.97	0.4	Nitrobenzene	0.102	1.20	0.014
					0.151	1.20	0.014

TABLE IV

Dissociation constants (K_b) of weak bases obtained by different methods

Base	Titration		Conductivity, ref. 9	Cryoscopy, ref. 10	Spectroscopy	
	$H_2S_2O_7$	$HB(HSO_4)_4$			Ref. 7	Ref. 11
p-Nitrotoluene	—	0.095	0.095	0.094	0.100	0.077
o-Nitrotoluene	—	0.071	0.067	0.062	—	—
m-Nitrotoluene	0.023	0.017	0.023	0.020	0.019	0.024
Nitrobenzene	0.013	0.009	0.010	0.011	0.010	0.009

shows that $HB(HSO_4)_4$ is a moderately strong acid but is not completely ionized as had been concluded previously (5) from freezing point and conductivity measurements on solutions of boric acid in sulphuric acid which contain the electrolyte $H_3O.B(HSO_4)_4$. However, the degree of solvolysis of this electrolyte implied by $K_a = 0.3$ is too small to have been detected in these measurements.

$HB(HSO_4)_4$, $H_2S_2O_7$ – *Weak Base Reactions*

Acid – weak base reactions were studied using the acids $H_2S_2O_7$ and $HB(HSO_4)_4$ and the weak bases *o*-, *m*-, and *p*-nitrotoluenes and nitrobenzene. Of these four bases only *m*-nitrotoluene and nitrobenzene could be successfully titrated with $H_2S_2O_7$ as the other two sulphonated at an appreciable rate. All four bases were quite stable in $HB(HSO_4)_4$ solutions. This would, incidentally, seem to confirm the suggestion (8) that sulphur trioxide which is present in appreciable concentration in $H_2S_2O_7$, but not in $HB(HSO_4)_4$ solutions, is the sulphonating agent rather than $H_3SO_4^+$, which is present in both solutions in appreciable concentration. The results of some of the conductivity measurements are shown in Fig. 4 and the observed end points, i.e. values of r_{min}, are given in Table III together with the values of K_b calculated by means of equations [20] and [22]. These values of K_b are in good agreement with those obtained by other methods, which are given in Table IV for comparison. This agreement confirms the reliability of this "conductimetric titration" method of determining the strengths of weak bases in the sulphuric acid solvent system.

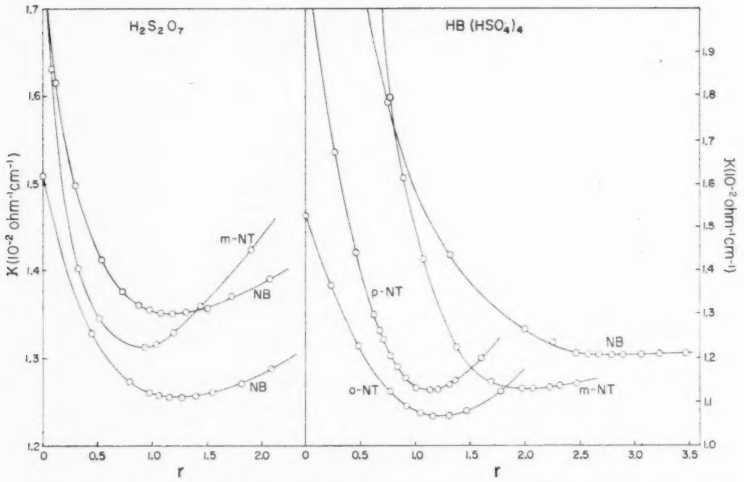

FIG. 4. Conductimetric titrations of weak bases with the acids $H_2S_2O_7$ and $HB(HSO_4)_4$. NB, nitrobenzene; *o*-NT, *o*-nitrotoluene; *m*-NT, *m*-nitrotoluene; *p*-NT, *p*-nitrotoluene.

EXPERIMENTAL

The apparatus and general procedure for making the electrical conductivity measurements have been described elsewhere (12).

Sulphuric acid was prepared as previously described (12) and the composition was adjusted until it had the conductivity corresponding to the composition 100% H_2SO_4 ($\kappa = 0.010439$ ohm^{-1} cm^{-1}) (12). Solutions of disulphuric acid were prepared by distilling sulphur trioxide from fuming sulphuric acid into "AnalaR" sulphuric acid. Solutions of $HB(HSO_4)_4$ were prepared by adding boric acid to an oleum, whose composition was known from its conductivity, in accordance with the stoichiometric equation

$$H_3BO_3 + 3H_2S_2O_7 = HB(HSO_4)_4 + 2H_2SO_4.$$

"AnalaR" grade potassium sulphate was dried at 120°. Benzoic acid, m.p. 121 to 122° C, was recrystallized from water and dried at 100°. *p*-Nitrotoluene, recrystallized from alcohol, had a melting point of 54°. Liquid solutes were dried and redistilled: nitrobenzene, b.p. 209°, $n_D^{25} = 1.550$; *o*-nitrotoluene, b.p. 218°, $n_D^{25} = 1.543$; *m*-nitrotoluene, b.p. 228°, $n_D^{25} = 1.544$.

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PROTON MAGNETIC RESONANCE SHIFTS FOR THE H_2O–SO_3 SYSTEM AND SOME ELECTROLYTE SOLUTIONS IN SULPHURIC ACID[1]

R. J. GILLESPIE[2] AND R. F. M. WHITE

ABSTRACT

Proton magnetic resonance shifts have been measured for the H_2O–SO_3 system. The results have been used to calculate the degree of ionization of water in solutions containing excess sulphuric acid. The shifts produced by a number of electrolytes in sulphuric acid have also been measured. The results are interpreted in terms of solute–solvent interactions, particularly polarization and hydrogen bonding, and are compared with the results of similar previous measurements on aqueous solutions.

INTRODUCTION

There have been two previous studies of the shift of the proton nuclear magnetic resonance in the H_2O–SO_3 system (1, 2). The present measurements on this system were undertaken as a preliminary to a more general study of the effects of both electrolytes and non-electrolytes on the proton resonance of sulphuric acid. Some differences were found between our results for the H_2O–SO_3 system and those of the earlier workers and so these results are given and discussed in some detail.

It has been found that electrolytes cause appreciable shifts in the proton resonance of water and this has been attributed to polarization and structure-breaking effects (3). Both cations and anions polarize water in such a manner that the electron density in the region of the protons is reduced and therefore resonance occurs at lower applied fields. The solvation of ions in general causes some structure-breaking in the solvent since the orientation of solvent molecules around ions is in general not the same as that in the bulk solvent and consequently a number of hydrogen bonds between the molecules in the solvation layer and those in the bulk of the solvent are broken. It has been established that generally the result of breaking hydrogen bonds is to shield the proton more effectively from the external field, thus shifting the resonance to higher applied fields (4). As ion–solvent interaction in sulphuric acid has been fairly extensively studied (5) it was of interest to see if proton shifts in sulphuric acid could be interpreted in a similar manner.

EXPERIMENTAL

The proton resonance shifts relative to pure water, $\delta = 10^6(H-H_{H_2O})/H_{H_2O}$, were measured with a Varian V4300B high-resolution spectrometer, operating at 40 Mc. Spectra were calibrated by the side-band technique using a Muirhead–Wigan L. F. Decade Oscillator. In preliminary experiments samples were sealed in 3-mm O. D. pyrex tubes, which were inserted into 5-mm O. D. tubes (2). The narrow annular space between the tubes contained distilled water which provided the reference signal. The results obtained using this type of sample tube showed a small scatter, probably due to tube imperfections. In order to reduce these effects the measurements were repeated using a different type of sample container. A small ground glass cone and socket joint was sealed to the end of a 5-mm O. D. sample tube. A reference capillary of water was sealed to the cone so that when it was inserted in the socket the capillary was centrally located in the sample tube. Scratch marks on the sample and reference tube ensured that they could

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always be put together in the same relative orientation. Although this technique does not completely eliminate effects due to tube imperfections any such effects will be almost constant in magnitude for all samples. Only results obtained using the second type of sample tube are reported; the results obtained from the first experiments agree with the reported results except for a somewhat greater scatter.

The diamagnetic susceptibilities of the solutions were measured by means of a Gouy balance (6). The apparatus available was designed for measurements on paramagnetic solids, and when used for diamagnetic solutions the accuracy obtainable was not better than 1%, mainly due to the small size of the sample. The measured susceptibilities and values calculated from Pascal's constants and the Wiedmann mixture law (6) indicate that the susceptibility correction (7)

$$g = (2\pi/3)\,(-0.721\times10^{-6}-\chi)\times10^{6},$$

where χ is the volume susceptibility of the solution, is almost negligible. As we give a largely qualitative interpretation of the observed shifts, susceptibility corrections have not in general been made except where stated. Sulphuric acid was prepared as previously described (8). Solutions containing small amounts of excess SO_3 or H_2O were analyzed by direct determination of their electrical conductivities (8) while those containing larger amounts of SO_3 or H_2O were analyzed by adding a known weight of the solution to a weighed amount of acid of known conductivity and measuring the change in conductivity.

RESULTS AND DISCUSSION

The Water – Sulphur Trioxide System

The results of our measurements on the H_2O–SO_3 system are given in Table I and Fig. 1.

TABLE I

Proton chemical shifts in the H_2O–SO_3 system

Solutions of composition between H_2O and H_2SO_4				Solutions of composition between H_2SO_4 and SO_3	
Mole fraction H_2O	δ (p.p.m.)	Mole fraction H_2O	δ (p.p.m.)	Mole fraction SO_3	δ (p.p.m.)
1.00	0.00	0.164	−6.39	0.0076	−5.95
0.913	−2.10	0.151	−6.36	0.013	−5.95
0.840	−3.40	0.149	−6.36	0.015	−5.95
0.685	−5.24	0.141	−6.34	0.022	−5.97
0.626	−5.78	0.091	−6.25	0.134	−5.98
0.510	−6.38	0.071	−6.19	0.207	−5.97
0.358	−6.57	0.045	−6.12	0.298	−5.96
0.323	−6.57	0.028	−6.03	0.411	−5.87
0.236	−6.48	0.015	−5.97	0.481	−5.79
0.197	−6.42	0.014	−5.98	0.509	−5.73
0.193	−6.43	0.006	−5.95	0.578	−5.62
		0.000	−5.95	0.723	−5.38

They agree well with Hood and Reilly's uncorrected results (2) up to about $x_{H_2SO_4} = 0.7$ but deviate from theirs at higher sulphuric acid concentrations. Although Gutowsky and Saika (1) only gave their results in graphical form they also appear to agree quite well with ours except at the highest sulphuric acid concentrations. Thus Hood and Reilly give $\delta_{H_2SO_4}$ (corrected) = −6.15 and $\delta_{H_2SO_4}$ (uncorrected) = −6.40, Gutowsky and Saika give $\delta_{H_2SO_4}$ (uncorrected) = −6.1, while we find $\delta_{H_2SO_4}$ (corrected) = −5.90 and $\delta_{H_2SO_4}$ (uncorrected) = −5.95. The earlier workers used reagent grade sulphuric acid for the

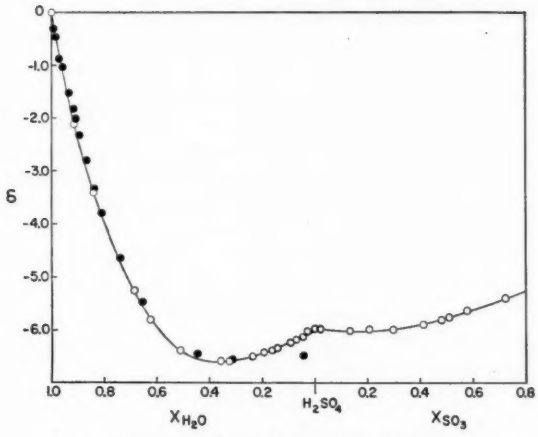

FIG. 1. Proton chemical shifts in the H_2O–SO_3 system.
○ This work; ● Hood and Reilly (2).

preparation of their solutions and determined their compositions by chemical analysis, while we have determined the composition of all our solutions by accurate conductivity measurements. It is likely that the differences in the uncorrected values of $\delta_{H_2SO_4}$ are merely a reflection of the fact that the composition of the more concentrated sulphuric acid solutions used by the earlier workers was not known with sufficient accuracy. Measurements for only two concentrations of oleum have been reported previously (1) and the same shift as for sulphuric acid was found in both cases. Our measurements on oleum solutions extend and essentially confirm these observations, showing that there is actually a very slight decrease in the shift in dilute oleum followed by a very flat minimum and a slight increase as the SO_3 concentration is increased up to and beyond the composition $H_2S_2O_7$.

Detailed analysis of the results is difficult because of the number of different species present in these solutions. We consider only the shifts in the region $x_{H_2SO_4} = 0.5$ to 1.0. Addition of water to sulphuric acid causes a shift in the proton resonance, which varies linearly with mole fraction up to approximately $x_{H_2SO_4} = 0.9$ (Figs. 1 and 3).

If we assume that the reaction that occurs when water is added to sulphuric acid can be represented simply by the equation

$$H_2O + H_2SO_4 = H_3O^+.HSO_4^-$$

then if x is the stoichiometric mole fraction of the water and α is the degree of ionization of the water the mole fractions of the species actually present in the solution are

$$N_{H_2O} = \frac{(1-\alpha)x}{1-\alpha x}, \qquad N_{H_2SO_4} = \frac{1-(1+\alpha)x}{1-\alpha x}, \qquad N_{H_3O.HSO_4} = \frac{\alpha x}{1-\alpha x}.$$

The corresponding proton fractions (i.e. the fractions of the total number of protons that are associated with any one species) are

$$p_{H_2O} = (1-\alpha)x, \qquad p_{H_2SO_4} = 1-(1+\alpha)x, \qquad p_{H_3O.HSO_4} = 2\alpha x.$$

Consequently the observed shift is given by

$$\delta = \delta_{H_2SO_4} - (1+\alpha)x\,\delta_{H_2SO_4} + 2\alpha x\,\delta_{H_3O.HSO_4}.$$

If the ionization of water is complete and $\alpha = 1$ then

$$\delta = \delta_{H_2SO_4} - 2x(\delta_{H_3O.HSO_4} - \delta_{H_2SO_4})$$

and we see that δ is a linear function of the mole fraction x. Experimentally this is found to be the case from $x = 0$ up to approximately $x = 0.10$. This indicates therefore that water is fully ionized over this range and by extrapolating this linear portion of the plot of δ against x to $x = 0.5$ we find $\delta_{H_3O.HSO_4} = -7.4$. At concentrations of water greater than $x = 0.10$ the curvature of the plot indicates that water is incompletely ionized. The degree of ionization α is given by

$$\alpha = \frac{\delta - \delta_{H_2SO_4}(1-x)}{x(2\delta_{H_3O.HSO_4} - \delta_{H_2SO_4})}.$$

Values of α obtained in this way are given in Table II and in Fig. 2 they are compared

TABLE II

Degree of ionization of water in H_2SO_4

Mole fraction H_2O	δ	α
0.5	−6.40	0.77
0.4	−6.56	0.84
0.3	−6.55	0.89
0.2	−6.46	0.96
0.1	−6.27	1.00

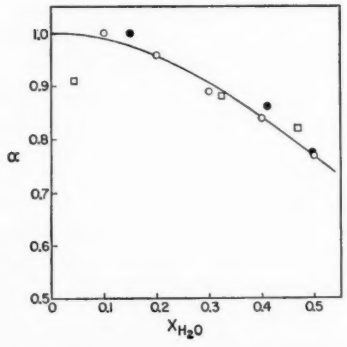

FIG. 2. Degree of ionization of water in sulphuric acid. ○ This work; ● Young, Maranville, and Smith (9); □ Hood and Reilly (2).

with values calculated from Hood and Reilly's results (1) and with values calculated from the Raman spectral data of Young *et al.* (9). In view of the assumptions involved in both methods the agreement between our results and those of Young *et al.* seems reasonable. The agreement with Hood and Reilly's results is also satisfactory, except at their lowest water concentration.

It is difficult to separate $\delta_{H_3O.HSO_4}$ into individual values for H_3O^+ and HSO_4^- on the basis of the results for the H_2O–SO_3 system alone. Hood and Reilly attempted to do this in the following way; they obtained expressions for two quantities symbolized as s_4 and

s_5, which are functions of the shifts $\delta_{H_3O.HSO_4}$ and $\delta_{H_2SO_4}$. The quantity s_4 was determined from the measured shifts in the region $x = 0$ to 0.2 while s_5 was estimated from the measured shifts in the region $x = 0.5$ to 1.0 using the second dissociation constant of sulphuric acid obtained from e.m.f. and conductivity data and the Raman spectral measurements of Young *et al.* (9) for this region. From these values of s_4 and s_5 they obtained $\delta_{H_3O^+} = -13.1$ and hence from their results it may be calculated that $\delta_{HSO_4^-} = 10.8$. Their calculations assume that the chemical shifts of the various species in the solution are independent of the composition of the solution over the whole range from water to sulphuric acid. We believe that this assumption is not justified. Considerable difficulties arise in interpreting the shifts of metal hydrogen sulphate solutions if the value $\delta_{HSO_4^-} = 10.8$ is accepted; as is shown in the following section, a negative value is more probable. Our calculations assume that the shifts produced by H_3O^+ and HSO_4^- are constant over the composition range $x = 0$ to 0.5, which is somewhat better than Hood and Reilly's assumption but is still only an approximation.

Metal Hydrogen Sulphate Solutions

The results of the measurements of the shifts produced by a number of metal hydrogen sulphates in sulphuric acid are given in Table III and Fig. 3. If x is the stoichiometric mole fraction of the metal hydrogen sulphate the proton fraction for a univalent metal hydrogen sulphate is $p_{MHSO_4} = x/(2-x)$ and for a divalent metal hydrogen sulphate it is

TABLE III

Proton chemical shifts of hydrogen sulphates in sulphuric acid

$H_3O.HSO_4$			$NaHSO_4$			$KHSO_4$		
x	p	δ	x	p	δ	x	p	δ
0.006	0.012	−5.95	0.0388	0.0198	−5.98	0.0505	0.026	−6.04
0.014	0.028	−5.98	0.1211	0.0645	−6.18	0.1128	0.0598	−6.22
0.015	0.030	−5.97	0.1740	0.0951	−6.31	0.1661	0.0907	−6.41
0.028	0.056	−6.03	0.2026	0.1125	−6.39	0.2464	0.1405	−6.67
0.045	0.090	−6.12						
0.071	0.142	−6.19						
0.091	0.182	−6.25						

$TlHSO_4$			NH_4HSO_4			
x	p	δ	x	p	δ (H_2SO_4 protons)	δ (NH_4^+ protons)
0.01644	0.00828	−6.04	0.0410	0.0209	−6.02	—
0.06309	0.0325	−6.21	0.1084	0.0573	−6.21	−1.13
0.08498	0.0444	−6.34	0.1232	0.0655	−6.25	−1.14
0.1440	0.0775	−6.57	0.1413	0.0760	−6.31	−1.15
			0.2331	0.1317	−6.62	−1.23
			0.2986	0.1758	−6.82	−1.28

$Ba(HSO_4)_2$		Acetone $(CH_3)_2COH.HSO_4$		
$x = p$	δ	$x = p$	δ (H_2SO_4 protons)	δ (CH_3 protons)
0.0134	−6.01	0.1753	−6.79	1.86
0.0239	−6.09	0.2439	−7.21	1.91
0.0314	−6.14	0.2616	−7.25	2.00
0.0386	−6.20			
0.0509	−6.34			

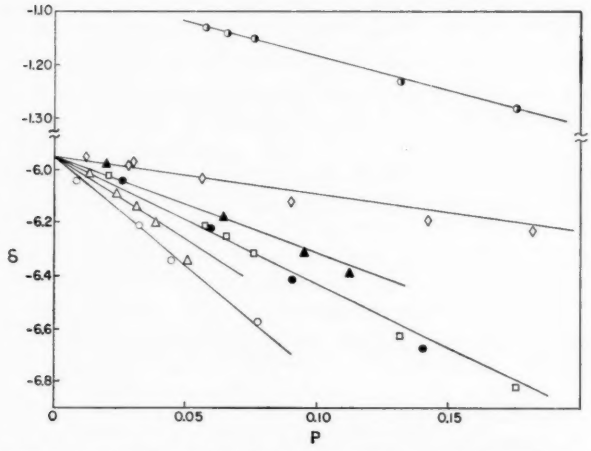

FIG. 3. Proton chemical shifts for hydrogen sulphates in sulphuric acid.
◑ HN_4; ◇ $H_3O.HSO_4$; ▲ $NaHSO_4$; ● $KHSO_4$;
□ $NH_4.HSO_4$; △ $Ba(HSO_4)_2$; ○ $TlHSO_4$.

$p_{M(HSO_4)_2} = x$. The proton shift would be expected to be the same linear function of the proton fraction p for all metal hydrogen sulphates:

$$\delta = \delta_{H_2SO_4} + p(\delta_{HSO_4^-} - \delta_{H_2SO_4}).$$

Figure 3 shows that the various metal hydrogen sulphates give linear, but nevertheless different, plots. It must be concluded that the metal ion also causes a shift in the proton resonance of the solvent in addition to that produced by the hydrogen sulphate ion. It has been found that electrolytes, even if they do not themselves contain any protons, cause appreciable shifts in the proton resonance of water (3); this has been attributed to polarization of the solvent molecules and breaking-up of the hydrogen-bonded structure of the solvent by the ions of the electrolyte. It has been shown that in aqueous solutions the polarizing effect of a cation is smaller than that of an anion of comparable size (3). This difference arises from the fact that a water molecule is oriented by a cation so that its hydrogen atoms are screened from the polarizing field by the oxygen atom. This screening effect would be considerably greater in sulphuric acid solutions because of the greater size of the sulphuric acid molecule compared with a water molecule. In sulphuric acid solutions the polarization effect of cations is therefore likely to be relatively small and the effect of cations on the proton resonance can most probably be attributed mainly to their structure-breaking effect on the solvent. The greater the extent of solvation of an ion the greater will be the number of hydrogen bonds broken and the more the resonance will shift to higher fields. Other measurements (5) have shown the extent of solvation of cations in sulphuric acid to be in the order

$$Na > NH_4 \sim K > Tl$$

and this is indeed the order of the shift of the proton resonance to high field (compared with Tl) in solutions of their hydrogen sulphates. Although the barium ion is more highly solvated than any of the univalent metal ions we have mentioned, solutions of barium di(hydrogen sulphate) give a smaller shift to high field (compared with Tl)

than solutions of sodium hydrogen sulphate. This can be attributed to the fact that at a given value of p the concentration of barium ion is only half that of a univalent metal ion and also to the fact that the dipositive charge of the barium ion will give rise to a much larger polarization effect, which may cause a significant shift of the proton resonance to low field.

Since the extent of solvation of the thallium ion is probably very small (5) it is reasonable to assume that its effect on the proton resonance is also small. An approximate value for the hydrogen sulphate ion shift may therefore be obtained by assuming that the shift produced by the thallium ion is in fact zero and that the whole of the shift observed for $TlHSO_4$ is due to the HSO_4^- ion. Hence we find $\delta_{HSO_4^-} = -13.9$. The positive value for $\delta_{HSO_4^-}$ obtained from the equations used by Hood and Reilly seems much less likely since in order to account for the observed shifts of metal hydrogen sulphates it would be necessary to attribute large low-field shifts to all metal ions, which in turn would suggest that all such metal ions produce very large polarization effects and this is rather unlikely. Moreover, if polarization of the solvent were the most important effect of the cations in these solutions it would be expected that the smallest ions would produce the greatest effects and consequently the order of the shifts of the proton resonance by different cations would be the reverse of that observed.

The large shift to low field produced by HSO_4^- is of interest. The electron density around the proton in an isolated HSO_4^- ion must be greater than that around the two protons in a sulphuric acid molecule and such an isolated HSO_4^- ion would be expected to have a proton resonance at a higher field than H_2SO_4. The shift to low field that is observed must be due to strong interaction with the solvent. Although the charge on HSO_4^- is rather dispersed it may nevertheless polarize the OH bonds of nearby sulphuric acid molecules since the hydrogen atoms will be oriented towards the HSO_4^- ion. This will cause a shift in the proton resonance of the solution to low field. The sulphuric acid molecules surrounding an HSO_4^- ion will be bound to it by hydrogen bonds and since the ion fits easily into the structure of sulphuric acid it may well cause no over-all bond-breaking but might well cause some increase in the total extent of hydrogen bonding in the solution, which would also cause a shift in the proton resonance to low field.

The shift produced by $H_3O.HSO_4$ is smaller than that of the metal hydrogen sulphates investigated and this may be attributed to the shift of H_3O^+ being smaller than that of H_2SO_4. We have that

$$\delta_{H_3O.HSO_4} = \delta_{HSO_4^-}/4 + 3\delta_{H_3O^+}/4$$

and using the values $\delta_{HSO_4^-} = -13.9$ and $\delta_{H_3O.HSO_4} = -7.4$ we find $\delta_{H_3O^+} = -5.2$. These values differ very considerably from those deduced by Hood and Reilly and they are probably only significant in systems containing excess H_2SO_4. It seems probable that the shifts produced by ions such as H_3O^+ and HSO_4^- which can hydrogen bond with the solvent change appreciably with the composition of the solvent. In order to account for the shift in the proton resonance caused by the addition of KNO_3 to HNO_3, Happe and Whittaker (10) have postulated the existence of the definite solvated species $NO_3^-.2HNO_3$ and $NO_3^-.HNO_3$ each producing its own shift and the latter gradually replacing the former with increase in the concentration of KNO_3. Similarly in predominately aqueous solutions of sulphuric acid it is very probable that the H_3O^+ ion exists in a solvated form such as $H_3O^+(H_2O)_3$ and as the concentration of acid is increased this is replaced by less hydrated complexes such as $H_3O^+(H_2O)_2$ and $H_3O^+(H_2O)$ and in strongly acid media by species such as $(H_3O^+)(H_2SO_4)_2$ (5, 11). It seems likely that the relatively

strong hydrogen-bonding in these complexes will cause each of them to have their own characteristic proton shifts. Unfortunately, if this is the case it makes any detailed analysis of this aqueous portion of the H_2O–SO_3 system almost impossible at the present time.

TABLE IV
Proton chemical shifts in HX–MX systems

Reference	HX–MX	Concn.	δ (ref. HX)
This work	H_2SO_4–$KHSO_4$	$p = 0.1$	−0.50
Gutowsky and Saika (1)	H_2O–KOH	$p = 0.1$	−1.00
Happe and Whittaker(10)	HNO_3–KNO_3	$x = 0.1$	−1.28
Unpublished (12)	HSO_3F–KSO_3F	$x = 0.1$	−1.38

Table IV shows the shifts in the proton resonance produced in different systems of the type HX–MX. All the shifts are negative, that is towards low fields, and they are all of the same magnitude. It seems reasonable to conclude that in each case X^- dissolved in HX produces a shift in the proton resonance to low field for the same reasons as given above for the HSO_4^-–H_2SO_4 system. The shifts are smallest in the cases of H_2O and H_2SO_4, which is consistent with the fact that in addition to the effects of interaction between the anions and the solvent the anions OH^- and HSO_4^- have their own intrinsic shifts, which must be to high field thus causing the over-all shift to low field to be smaller than the shift due to the anion–solvent interactions alone.

The Ammonium Ion

The proton resonance of the ammonium ion was observed as a separate peak indicating that these protons do not undergo rapid exchange with those of the solvent. The peak is a triplet due to spin–spin coupling with the nitrogen nucleus and shifts were measured from the center of this triplet which occurs on the low-field side of water and shifts to lower fields with increasing concentration (Table III and Fig. 3). This shift with increasing concentration can be plausibly attributed to an increased strength of the hydrogen bonds between the ammonium ion and the solvent as a consequence of the increasing basicity of the solvent with increasing concentration of HSO_4^-.

Acetone

Acetone produces a shift in the sulphuric acid resonance towards low field which varies linearly with mole fraction and which is of comparable magnitude to that produced by the metal hydrogen sulphates (Table III). Acetone is fully ionized (13) according to the equation

$$(CH_3)_2CO + H_2SO_4 = (CH_3)_2COH^+ + HSO_4^-$$

and if x is the mole fraction of $(CH_3)_2COH.HSO_4$ and the proton fraction is p then $p = x$. No detailed interpretation is possible at present because we cannot separate the intrinsic shift of $(CH_3)_2COH^+$ from the effect of its interaction with the solvent.

It is interesting that there is an appreciable shift in the CH_3 proton resonance to high field with increasing concentration of acetone. This could be due to a direct interaction of the CH_3 protons with the solvent, i.e. hydrogen-bonding or polarization of the C—H bond electrons by the negatively charged oxygen atoms of neighboring sulphuric acid molecules. Both these effects would lead to a shift in the proton resonance to low field with increasing concentration of hydrogen sulphate ion produced by the ionization of

the acetone. The observed shift is, however, to high field; the direct interaction of the CH_3 group with the solvent must therefore be small and the observed shift must be attributed to an indirect effect of the interaction of some other part of the molecule with the solvent. The protonated carbonyl group will be hydrogen-bonded to the solvent to an extent that will increase with the concentration of hydrogen sulphate ion and the positive charge on the oxygen atom will diminish correspondingly. This positively charged oxygen atom will polarize the C—H bond electrons of the CH_3 group and this polarization will therefore diminish with increasing concentration of hydrogen sulphate ion produced by the ionization of the acetone causing a shift in the resonance of these protons to high field.

The Fluorosulphuric Acid – Sulphuric Acid System

If the stoichiometric mole fraction of fluorosulphuric acid is x the corresponding proton fraction is $p_{HSO_3F} = x/(2-x)$. A plot of the shift of the proton resonance (Table V) against p_{HSO_3F} is linear (Fig. 4) as would be expected if there were no ionization or other interaction between the fluorosulphuric acid and the sulphuric acid. It has been found that HSO_3F is very slightly ionized in H_2SO_4 (14) but no effect of this is seen in the proton shifts.

The Chlorosulphuric Acid – Sulphuric Acid System

The proton resonance is independent of the composition of the system (Fig. 4 and Table V) and is the same as that for sulphuric acid itself. Presumably there is no ionization nor any appreciable change in hydrogen-bond strengths.

TABLE V

Proton chemical shifts in acid – sulphuric acid systems

Electrolyte	x	p	δ
$HB(HSO_4)_4$	0.037	0.122	−6.11
	0.118	0.311	−6.47
	0.264	0.662	−6.84
	0.420	0.901	−7.09
$HFSO_3$	0.182	0.100	−5.88
	0.287	0.168	−5.86
	0.311	0.184	−5.86
	0.639	0.470	−5.65
	0.767	0.624	−5.54
	1.000	1.000	−5.24
$HClSO_3$	0.216	0.121	−5.95
	0.417	0.263	−5.95
	0.735	0.581	−5.95
	1.000	1.000	−5.95

The Tetra(hydrogensulphato) Boric Acid – Sulphuric Acid System

Tetra(hydrogensulphato) boric acid, $HB(HSO_4)_4$, has been shown to ionize in sulphuric acid solvent as a moderately strong acid (14, 15, 16).

$$HB(HSO_4)_4 + H_2SO_4 = H_3SO_4^+ + B(HSO_4)_4^-$$

If the stoichiometric mole fraction of $HB(HSO_4)_4$ is x then the mole fractions of the species actually present in solution are

$$N_{H_2SO_4.B(HSO_4)_4} = x/(1-x) \text{ and } N_{H_2SO_4} = (1-2x)/(1-x)$$

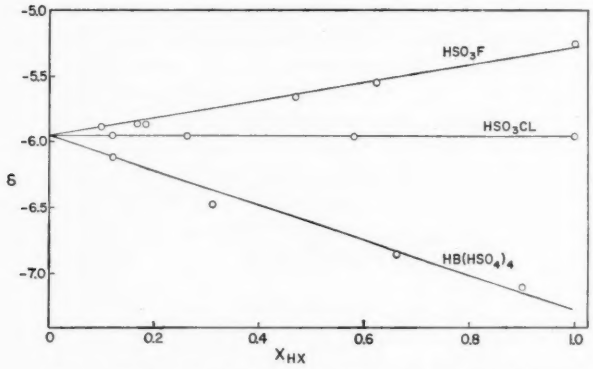

FIG. 4. Proton chemical shifts in HX – sulphuric acid systems.

and the corresponding proton fractions are

$$p_{H_2SO_4.B(HSO_4)_4} = 7x/(2+3x), \quad p_{H_2SO_4} = (2-4x)/(2+3x)$$

and the proton shift would be expected to be a linear function of $p_{H_2SO_4.B(HSO_4)_4}$:

$$\delta = \delta_{H_2SO_4} + p_{H_2SO_4.B(HSO_4)_4}\{\delta_{H_2SO_4.B(HSO_4)_4} - \delta_{H_2SO_4}\}.$$

The observed shift (Fig. 4 and Table V) is in fact a linear function of $p_{H_2SO_4.B(HSO_4)_4}$ and is smaller than that due to the metal hydrogen sulphates. There is not sufficient information on this system available at present to enable a detailed analysis of the observed shifts to be made.

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THE PRESENCE OF MONOTHIOCARBONATE SUBSTITUENTS IN CELLULOSE XANTHATES[1]

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ABSTRACT

The presence of monothiocarbonate substituents in viscose has been demonstrated by the infrared spectra of the gases evolved during the acid hydrolysis of cellulose xanthate salts and viscose. A comparison of the results of various methods of sulphur determination and the results obtained by ultraviolet spectroscopy has been made. The results of investigations on sodium cellulose monothiocarbonate preparations support the concept of monothiocarbonate substitution.

INTRODUCTION

Sodium cellulose xanthate as normally prepared commercially from alkali cellulose and carbon disulphide is accompanied by numerous by-products. The composition of this mixture and of its solution in aqueous sodium hydroxide (viscose) changes as a result of complex hydrolytic and oxidative reactions. It is well established that the proportion of trithiocarbonate, the main by-product, increases as a result of the hydrolysis of sodium cellulose xanthate and that hydrolysis and oxidation of the trithiocarbonate lead to sulphide, perthiocarbonate, and thiosulphate (1, 2, 3, 4). The presence of other components, often suggested to account for the complex reactions, is less clearly established. Mono- and dithio-carbonate ions (5, 6) have been invoked as intermediate stages in the xanthation and side reactions, and cellulose monothiocarbonate ions can be postulated as a reaction product of monothiocarbonate ions and cellulose and as a hydrolysis product of sodium cellulose xanthate. Monothiocarbonates yield carbonyl sulphide on acid hydrolysis and the presence of this substance among the gaseous products of acidified viscose would provide substantial evidence for the existence of such derivatives.

RESULTS AND DISCUSSION

Infrared Spectra of Gases Evolved from Acidified Viscose

The presence of carbonyl sulphide in gaseous mixtures containing carbon disulphide, carbon dioxide, and hydrogen sulphide can be demonstrated by the infrared spectrum, and considerably less than 5% can be readily detected by the most sensitive band at 2030 cm^{-1} (Fig. 1).

The gaseous products of acidified viscose undoubtedly contained carbonyl sulphide (Fig. 2, Table I), and essentially the same relative proportion of carbonyl sulphide was obtained from impure viscose and from a solution of sodium cellulose xanthate purified from small anions by passage through an anion exchange column (7). This observation strongly suggests that much of the carbonyl sulphide was chemically linked to the cellulose, a conclusion supported by the observation that the viscose impurities formed from alkali and carbon disulphide gave a much smaller proportion of carbonyl sulphide than freshly prepared viscose. The relative proportion of carbonyl sulphide was smaller for a "ripened" viscose, an observation in agreement with the known fact that cellulose monothiocarbonate hydrolyzes more rapidly than cellulose xanthate (8).

These results are in agreement with those recently reported by Philipp (9), who

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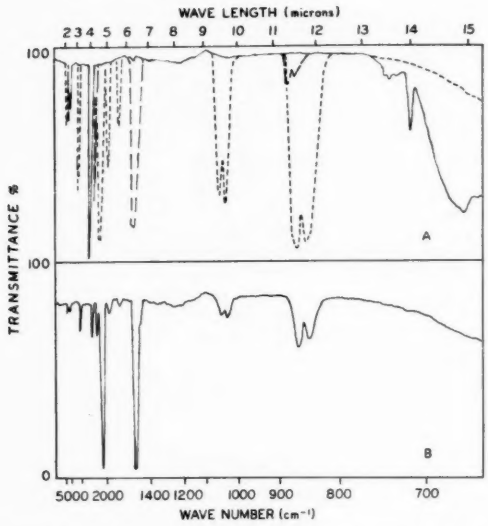

FIG. 1. Infrared spectra: (A) CO_2 ———, COS - - - -, CS_2 — — —; (B) mixture (76 ml), H_2S 88%, COS 4%, CS_2 8%.

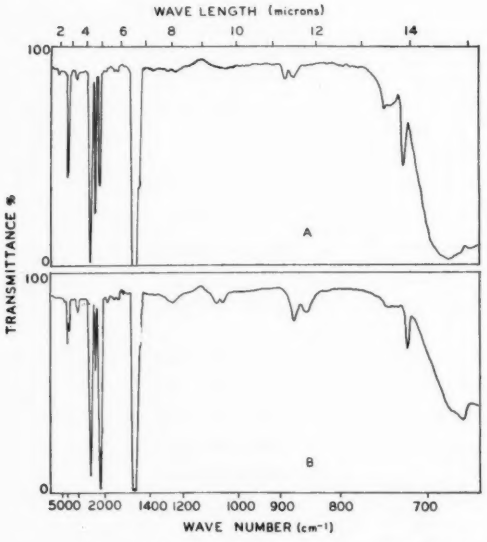

FIG. 2. Infrared spectra: gases from (A) purified "ripened" viscose (150 ml), (B) freshly prepared sodium cellulose xanthate (61 ml).

estimated the carbonyl sulphide evolved from viscose polarographically as ammonium monothiocarbamate. Very little reliance, however, can be placed on the quantitative data reported in this paper. Carbon disulphide, absorbed as ammonium dithiocarbamate, was found to seriously interfere in the polarographic estimations, and, under the conditions

TABLE I

Proportional band intensities of COS and CS_2 obtained on gases generated by hydrolysis

Hydrolysis gases from:	$\frac{(\log T_0/T \text{ (COS at 2030 cm}^{-1})}{(\log T_0/T \text{ (CS}_2 \text{ at 2179 cm}^{-1})}$
(1) Sodium cellulose xanthate	
(*a*) Freshly prepared	8.56
(*b*) Viscose (solution of (*a*) in 7% sodium hydroxide; 7% cellulose concentration)	3.23
(*c*) (*b*) Anion-exchanged concentrated	4.42
(*d*) (*b*) Anion-exchanged dilute	2.55
(*e*) After 2 days ripening at 8° C	0.78
(*f*) (*e*) Anion-exchanged concentrated	0.85
(*g*) (*e*) Anion-exchanged dilute	0.66
(2) Potassium cellulose xanthate Dissolved in 5% potassium hydroxide; approx. 7% cellulose concentration; anion-exchanged dilute	12.10
(3) NaOH + CS_2 (viscose by-products)	1.18

used to minimize its absorption, only a portion of the total carbonyl sulphide was absorbed. The necessary large corrections render the method inaccurate. In addition, cadmium acetate was used to remove hydrogen sulphide from gaseous mixtures, a reagent which we found to react to some extent with carbonyl sulphide. In contrast to our observation, Herrent and Inoff (10) concluded from indirect analytical methods of somewhat doubtful validity that the proportion of monothiocarbonate increased as viscose ripened.

A quantitative assessment of the proportion of carbonyl sulphide has not been attempted. The gaseous hydrolytic products of viscose could be completely freed from the precipitated cellulose only by prolonged boiling which, in presence of acid, is said to decompose carbonyl sulphide (9). Quantitative removal of the gases in a stream of nitrogen at room temperature and subsequent absorption in alcoholic potassium hydroxide or ammonia solutions was attempted, but a quantitative recovery of the carbonyl sulphide was not possible since these reagents gave by-products, sulphide and carbonate, in addition to monothiocarbonate and monothiocarbamate, respectively.

Comparison of Results of Various Sulphur Determination Methods

Xanthates in solution are commonly estimated by oxidation to sulphate and by iodometric titration (11). Both methods are said to be quantitative for sulphides and monothiocarbonates, but we have not been able to confirm a quantitative reaction, according to equation 1, of iodine and potassium ethyl monothiocarbonate.

$$2RO{-}COS^- + I_2 \rightarrow 2I^- + RO{-}CO{-}S{-}S{-}CO{-}OR \qquad [1]$$

The presence of these substances in mixtures with xanthate, however, should be revealed by application of both methods: the assumption that sulphur is present only as xanthate would lead to higher calculated amounts of sulphur from the iodine method. In many sets of careful analyses of a variety of commercially prepared viscose samples, of which two typical examples are given (Table II, (*a*) and (*b*)), the amounts of sulphur estimated iodometrically in samples purified by ion exchange were not systematically different from those estimated as sulphate. Purified samples of potassium cellulose xanthates, however,

TABLE II

Comparison of results of various sulphur determination methods on sodium and potassium cellulose xanthates

		Not ion-exchanged			Ion-exchanged				
					Concentrated	Dilute			
		% Sulphur				% Sulphur			
						Total by:		By hydrolysis	
No.	Ripening time (days)	Total (by $BaSO_4$)	As CS_2^{2-}	Salt index	% Sulphur total (by $BaSO_4$)	$BaSO_4$	Iodine	H_2S	CS_2
		(*a*) Commercial viscose (sodium cellulose xanthate dissolved in 6% NaOH)							
1	—	2.31	0.36	—	—	1.82	1.82	0.02	1.75
2	—	1.90	0.22	11.2	—	1.48	1.46	—	—
3	—	1.73	1.36	—	—	1.45	1.45	—	—
4	—	2.06	0.32	9.7	—	1.44	1.57	0.05	1.45
5	—	1.78	0.30	11.5	—	1.42	1.43	—	—
6	—	1.86	0.60	4.0	—	1.27	1.20	0.03	1.15
		(*b*) Commercial viscose (series ripened at 8° C)							
1	—	1.83	0.29	10.5	1.39	1.44	1.39	—	—
2	1	1.85	0.34	9.7	1.35	1.36	1.29	—	—
3	2	1.82	0.39	8.9	1.31	1.36	1.30	0.03	1.30
4	3	1.81	0.42	8.1	1.29	1.30	1.29	—	—
5	4	1.80	0.45	7.4	1.25	1.25	1.23	0.02	1.27
6	7	1.79	0.56	5.6	1.16	1.13	1.12	0.01	1.19
7	8	1.81	0.57	5.1	1.09	1.12	1.11	0.01	1.06
8	9	1.79	0.63	4.7	1.03	1.09	1.09	—	1.02
9	10	1.79	0.64	4.0	—	1.02	1.03	—	—
		(*c*) Potassium cellulose xanthate (dissolved in 5% KOH)							
1	—	2.20	0.90	—	—	1.34	3.51	—	—
2	—	4.13	—	—	—	2.13	2.84	—	—
3	—	5.05	3.25	—	—	1.98	2.33	—	—
4	—	6.04	3.53	—	—	2.71	3.09	—	—
5	—	5.13	2.31	—	—	2.90	3.40	—	—
6	—	7.38	4.17	—	—	2.97	3.04	—	—
7	—	6.27	4.93	—	—	1.73	2.75	—	—
8	—	6.44	3.46	—	—	2.67	2.88	—	—

NOTE: All analytical data given are mean values from triplicate determinations.

showed large differences (Table II (*c*)). These samples, of higher degree of substitution, gave a much larger proportion of carbonyl sulphide on hydrolysis (Table I), and it is evident that the proportion of monothiocarbonate in commercial viscose is not sufficiently large to produce significant differences in the two analytical methods.

Ultraviolet Spectra of Various Cellulose Xanthates

In solution, absorption bands are shown by sodium cellulose xanthate at 382 and 302 mμ and by potassium ethyl xanthate at 380 and 300 mμ (12, 13). The extinction coefficients (Table III) are similar and appear to increase slightly with concentration.

The mean values for the molecular extinction coefficients of potassium ethyl xanthate and sodium cellulose xanthate in the 300–302 mμ region are 19,130 and 16,710 liters $mole^{-1}$ cm^{-1} respectively. Since monothiocarbonate does not absorb at 300 mμ, the extinction coefficient at 300 mμ of 1 mole of pure xanthate is therefore 19,130. Assuming that the lower extinction coefficient of sodium cellulose xanthate is due to the presence of monothiocarbonate, then the ratio of monothiocarbonate to xanthate groups in sodium cellulose xanthate would be

$$(19{,}130 - 16{,}710)/19{,}130 = 0.127.$$

TABLE III

Extinction coefficients of potassium ethyl xanthate and sodium cellulose xanthate in aqueous NaOH (pH 11.3)

	380 mμ		382 mμ			300 mμ		302 mμ	
	Sample					Sample			
	Et—O—CS_2—K		Cell—O—CS_2—Na			Et—O—CS_2—K		Cell—O—CS_2—Na	
Slit width, mm	c, moles/l. $\times 10^3$	ϵ, l./mole	c, moles/l. $\times 10^3$	ϵ, l./mole	Slit width, mm	c, moles/l. $\times 10^5$	ϵ, l./mole $\times 10^{-4}$	c, moles/l. $\times 10^5$	ϵ, l./mole $\times 10^{-4}$
0.125	1.951	60.48	3.152	58.07	0.087	0.975	1.754	1.051	1.625
					0.145		1.764		1.694
0.026	3.902	58.94	5.253	56.36	0.310				1.713
0.032		58.13		56.73					
0.060		58.18		57.11	0.087	1.951	1.835	2.101	1.594
					0.145		1.871		1.613
0.025	7.804	56.13	6.303	57.91	0.310		1.896		1.623
0.032		56.38		57.91					
0.040		56.51		58.22	0.087	3.902	1.881	3.152	1.618
0.060		56.66		58.54	0.145		1.907		1.644
0.112		57.02		58.54	0.310		1.922		1.648
0.026	11.710	56.64	8.404	57.12	0.093	5.853	1.865	4.202	1.618
0.038		56.72		57.71	0.145		1.931		1.642
0.115		56.06		57.82	0.310		1.961		1.654
0.025	15.610	56.00	10.510	57.69	0.087	7.804	1.961	5.253	1.618
0.038		56.51		58.07	0.142		1.986		1.647
0.115		56.64		58.26	0.310		2.063		1.656
0.025	19.510	56.30	12.600	57.35	0.087			6.303	1.623
0.038		56.36		57.59	0.145	0.981	1.856		1.653
0.112		56.38		58.29	0.310				1.662
0.025			14.710	58.47	0.087			8.404	1.648
0.038				59.02	0.145	1.961	1.882		1.725
0.112				59.84	0.310				1.731
0.025			16.810	56.40	0.087			10.510	1.713
0.038				57.18	0.145	2.941	1.904		1.866
0.112				57.71	0.310				1.884
0.025			18.910	57.00	0.310	3.921	1.887		
0.038				58.17	0.310	5.888	1.917		
0.112				58.49	0.310	7.844	2.014		
0.025			21.010	58.21					
0.038				59.02					
0.112				59.49					

If this value is now applied to the comparison of the percentage of sulphur estimated iodometrically and as barium sulphate the percentage of sulphur values calculated from the iodometric results should be considerably larger than those actually found (Table II). It is therefore not possible to consider the ultraviolet spectral data as conclusive evidence for the presence of monothiocarbonate substituents in the sodium cellulose xanthates investigated.

A comparison of the molecular extinction coefficients of sodium and potassium cellulose xanthate (Table IV) shows an even larger difference (e.g. 16,710 vs. 11,860). The fact that the percentage of sulphur values of potassium cellulose xanthate preparations calculated as xanthate from the iodometric method are consistently larger than those obtained by the barium sulphate method as well as the observed presence of large amounts

TABLE IV
Extinction coefficients of sodium cellulose xanthates and potassium cellulose xanthates

λ = 302 mμ, sodium cellulose xanthates				λ = 303 mμ, potassium cellulose xanthate		
Sample No.*	c,† moles/l. ×10⁵	At pH 10 ε, l./mole ×10⁻⁴	At pH 6 ε, l./mole ×10⁻⁴	Sample No.‡	c,‡ moles/l. ×10⁵	At pH 10 ε, l./mole ×10⁻⁴
(*b*)/1	2.52	1.49	1.49	(*c*)/1	2.10	1.03
	3.50	1.52	1.49		4.19	1.01
	6.76	1.54	1.52		6.29	1.04
2	2.13	1.49	1.51	2	3.32	1.14
	4.25	1.53	1.52		6.64	1.15
	6.38	1.54	1.56		9.97	1.24
3	2.13	1.41	1.45	3	2.69	1.45
	4.26	1.46	1.51		6.17	1.17
	6.39	1.50	1.52		9.26	1.22
4	2.09	1.41	1.44	4	4.23	1.04
	4.17	1.44	1.46		8.47	1.06
	6.26	1.48	1.48		12.70	1.10
5	1.95	1.39	1.44	5	4.52	1.39
	3.90	1.48	1.48		9.02	1.49
	5.85	1.52	1.50		13.55	1.34
6	1.78	1.58	1.59	6	4.63	1.17
	3.56	1.59	1.60		9.26	1.18
	5.34	1.62	1.61		13.46	1.27
7	1.74	1.54	1.57	7	3.45	1.17
	3.48	1.60	1.60		5.17	1.16
	5.23	1.61	1.63		10.08	0.98
8	1.71	1.47	1.47	8	4.20	1.20
	3.42	1.51	1.54		8.37	1.23
	5.14	1.58	1.55		12.59	1.23
9	1.59	1.45	1.59			
	3.19	1.49	1.50			
	4.80	1.54	1.55			

*Samples are from the sodium cellulose xanthate series mentioned in Table I, Section (*b*).
†Determined from sulphur as $BaSO_4$ (mean values).
‡Samples are from the potassium cellulose xanthates mentioned in Table I, Section (*c*).

of carbonyl sulphide in the gases generated from the hydrolysis are indicative of the presence of monothiocarbonate substituents.

The formation of the colored by-products in purified cellulose xanthate solutions may readily be detected by a shoulder at 325 mμ; higher concentrations increase the extinction coefficient at 382 mμ and eventually produce a slight shift to longer wavelengths. These criteria are more sensitive than the detection of the trithiocarbonate absorption band at 500 mμ. The band at 382 mμ, but not that at 302 mμ, is immediately altered at a pH of between 4 and 6. The maximum is first shifted to 390 mμ and then the absorbance rapidly decreases. This observation supports previous statements (15) that the bands at 302 and 382 mμ are a result of the excitation of the C=S bond and the complete xanthate group, respectively.

Investigations on Sodium Cellulose Monothiocarbonate

Sodium cellulose monothiocarbonate was prepared at a low temperature from alkali cellulose and carbonyl sulphide, which had been carefully purified until carbon disulphide could not be detected spectroscopically. The freshly prepared white products were

partially soluble in sodium hydroxide solution (7%), but one experiment gave a completely soluble product. About two thirds of the sulphur present was not retained by an anion exchange resin (Table V), and some carbonyl sulphide was regenerated on acid hydrolysis.

TABLE V
Comparison of results of various sulphur determination methods on sodium cellulose monothiocarbonate

	Not ion-exchanged				Ion-exchanged			
		% Sulphur				% Sulphur		
Sample No.	% Cell.	Total (by $BaSO_4$)	By hydrolysis H_2S	By hydrolysis COS	% Cell.	Total (by $BaSO_4$)	By iodine as monothiocarbonate	D.S.
1	16.44	5.29	—	—	12.10	1.81	2.85	0.19*
2	25.58	2.78	1.54	0.67	25.10	1.99	3.25	0.10*
3	12.15	3.00	—	—	10.90	2.00	2.44	0.23*
4	10.56	2.24	—	—	15.32	2.41	3.15	0.20†

*White, from pure COS.
†Prepared from COS containing CS_2; green.

In contrast to the results obtained using sodium cellulose xanthate, the iodometrically derived sulphur values, assuming pure monothiocarbonate, are higher than those derived from barium sulphate estimations, indicating appreciable decomposition to sulphide under the conditions of the purification and estimation.

The cellulose monothiocarbonate solutions showed no spectroscopic absorption bands in the visible and ultraviolet regions at 385 and 302 mμ and cellulose xanthate was therefore absent. Only one preparation of cellulose monothiocarbonate, obtained from impure carbonyl sulphide, contained any xanthate as shown by its spectrum (cf. Fig. 3). Potassium ethyl monothiocarbonate, however, similarly prepared from carbonyl sulphide

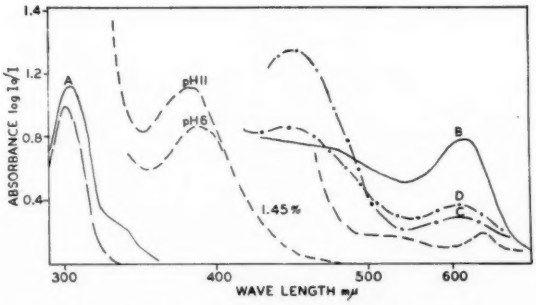

FIG. 3. Ultraviolet and visible absorption spectra of solutions of purified viscose (– – –); sodium cellulose monothiocarbonate (impure) (———): (A) in alkali, (B) acetone extract; aqueous sodium sulphide, carbon disulphide, and acetone (— · —): (C) 20 hours, (D) 26 hours.

(spectroscopically free of carbon disulphide) and potassium hydroxide in ethanol, was accompanied by detectable amounts of potassium ethyl xanthate.

In previous descriptions (14) of the reaction between alkali cellulose and carbonyl sulphide, it is stated that the product is not completely alkali soluble, but aside from some qualitative X-ray data, no analytical details are given. Philipp (8) describes the product

as a green solid. Our products were green only when the carbonyl sulphide contained carbon disulphide. Acetone produced a more intense green or blue color, which soon faded in air, and sulphur precipitated. Such behavior is reminiscent of the oxidation of hydrosulphides, said to produce disulphides responsible for the intense color (15). Similar colors are sometimes to be seen in the mother liquors from viscose precipitated with acetone, in reaction mixtures of sodium sulphide, carbon disulphide, and acetone, and in freshly prepared viscose samples as they emerge from a column of anion exchange resin. In all cases an absorption band at about 650 mμ was observed (Fig. 3).

Infrared Spectra of Solid Samples

The infrared spectra of impure and 96% pure potassium ethyl monothiocarbonate are shown in Figs. 4A and 4C. The spectrum of the fraction containing the impurities is shown in Fig. 4B. Monothiocarbonate is characterized by important absorption bands in

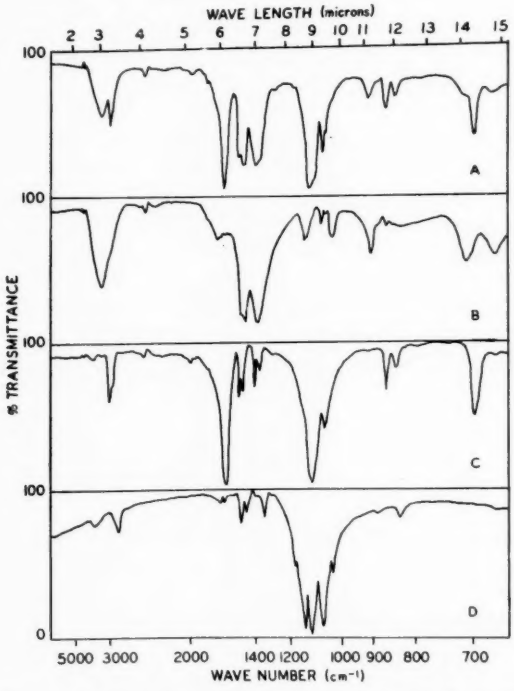

FIG. 4. Infrared spectra of potassium ethyl monothiocarbonate: (A) impure product, (B) insoluble fraction, (C) second recrystallization (96% pure). Infrared spectra of potassium ethyl xanthate (D).

the 1000–1200 cm^{-1} region, and at 1650 cm^{-1}. A less intense band occurs at 690 cm^{-1} and this is probably due to the C—S vibration. Comparison of the spectrum of potassium ethyl monothiocarbonate with that of potassium ethyl xanthate (Fig. 4D) shows that only the 1650 cm^{-1} band and to some extent the 690 cm^{-1} band can be used to detect monothiocarbonate groups in the presence of xanthate groups.

In the spectra of xanthates, intense absorption is always found in the 1060–1040 cm^{-1} region. At least two additional bands are usually found in the 1200–1100 cm^{-1} region. The

bands in the 1200–1000 cm^{-1} region are thought to be due to vibrations of the R—O—C part of the molecule.

The infrared absorption spectra of cellulose and alkali cellulose are shown in Figs. 5A and 5B. The nature of the bands in the cellulose spectrum have previously been discussed (16). The differences between the alkali cellulose spectrum and that of cellulose are mainly due to the presence of carbonate bands and to a general lowering of the cellulose bands due to changes in fine structure.

The infrared absorption spectrum of dried samples of sodium cellulose monothiocarbonate (Figs. 5C and 5D) revealed only the bands of cellulose and carbonate. Freshly

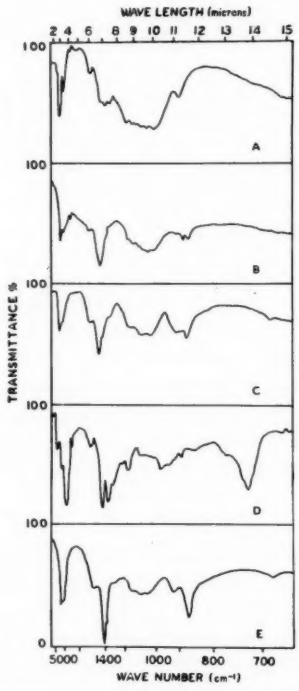

FIG. 5. Infrared spectra of cellulose (A), alkali cellulose (B), dried sodium cellulose monothiocarbonate (C, D), and freshly prepared wet sodium cellulose monothiocarbonate (E).

prepared wet samples were made into potassium bromide pellets with much difficulty (Fig. 5E). The main bands shown by this sample are those due to carbonate.

In an attempt to provide positive evidence for the presence of cellulose monothiocarbonate, the infrared spectra of potassium bromide pellets of samples of sodium cellulose xanthate precipitated from various viscose preparations (Fig. 6A) were examined as well as the corresponding S-methyl (Fig. 6B) and diethylacetamido (Fig. 6C) derivatives.

No bands analogous to those which can be ascribed to the monothiocarbonate are to be observed (Fig. 4). The degree of substitution of the cellulose monothiocarbonate samples was low, the bands are not pronounced and, as the amounts of monothiocarbonate were much less, it is not surprising that it escaped detection.

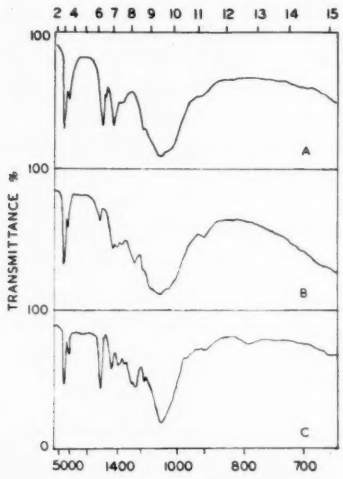

FIG. 6. Infrared spectra of cellulose xanthate precipitated from viscose (A), its S-methyl derivative (B), and its diethylacetamido derivative (C).

The precipitation of the cellulose xanthate salts by acetone or methanol at 0–5° C (17) and of the diethylacetamido derivative (18) from viscose by described methods was found to be accompanied by some loss of xanthate groups. The previous work showed no such loss. The two pronounced bands found for the sodium cellulose xanthate samples at 1563 and 1428 cm^{-1} were probably the result of impurities. Similar bands were found for some samples of the S-methyl derivative and these disappeared after successive extractions with water, acetone, and ether.

EXPERIMENTAL

Materials and Analytical Methods

Commercial samples of hydrogen sulphide (19), carbon disulphide, and carbon dioxide (20) were purified by established methods. Carbonyl sulphide was passed through sodium hydroxide solution (33%) and concentrated sulphuric acid and then liquefied by cooling with solid carbon dioxide and acetone. Two distillations through an externally cooled (−35° C) Vigreux column at a reflux ratio of about 3:1 were necessary to obtain carbonyl sulphide free of carbon disulphide. Triethylphosphine in nitrobenzene (1:10) also absorbed carbon disulphide from the impure mixture (21, 22), but some could still be detected in the infrared spectrum of the carbonyl sulphide.

Generation of Gases

One arm of a long U-tube (1 m × 2 cm diameter), separated from the other arm by a stopcock and containing the viscose (20 g), sodium cellulose xanthate or monothiocarbonate (5–10 g), was evacuated and phosphoric acid (50%) was admitted from the other arm. The heat necessary to free the gases from the cellulose was supplied externally. The reaction product of sodium hydroxide (8%, 160 ml) and carbon disulphide (25 ml) at room temperature (18 hours) was hydrolyzed similarly.

All gases were dried with phosphorus pentoxide and stored over mercury. Measured volumes in a 10-cm cell were examined in a Perkin–Elmer model 13 infrared spectrophotometer.

Preparation of Viscose

Viscose was prepared by standard commercial methods (23). All samples and solutions were kept at −5° to 0° C to minimize hydrolytic decomposition, and analyses were performed immediately after purification. Samples of viscose were allowed to "ripen" in a constant temperature bath at 8° C.

The total sulphur in viscose (10 g) diluted in sodium hydroxide solution (8%) at −5° C to 100 ml was estimated in aliquots (20 ml) by analytical method No. 2 below. The trithiocarbonate was estimated approximately from the extinction at 500 mμ of samples in 1-cm silica cells. (An extinction coefficient of 449 liters $mole^{-1}$ cm^{-1} was found using sodium trithiocarbonate solutions (16) prepared under nitrogen and estimating the trithiocarbonate content by method 1.)

Purification of Viscose

Amberlite anion exchange resin IRA 400 (8, 24) was used for purifying the viscose.

1. Viscose (100 g) was forced through the resin (200 g) by air pressure (30 lb in^{-2}) in an apparatus similar to that described by Dux (24).

2. Viscose (25 g) at 0° C was diluted with sodium hydroxide solution (8%, 250 ml) and passed through an externally cooled (−5° C) resin column (100 g, 2 cm diameter) and eluted to 1 liter. Dilution and elution with water instead of sodium hydroxide solution was found to cause a remarkable increase in elution time, particularly for cellulose xanthate of high xanthate substitution (D.S. = 0.7). The viscosity of cellulose xanthate solutions decreases with increasing alkali concentration (25).

Analytical Methods

1. Aliquots of purified viscose solution (100 ml) were oxidized by hydrogen peroxide (30%, 20 ml). After 10 minutes at room temperature, the solutions were diluted (100 ml), boiled to decompose the excess peroxide (20 minutes), cooled, made up accurately to 250 ml, and filtered. The sulphate in an aliquot (150 ml) was estimated gravimetrically as barium sulphate (26).

2. Similar aliquots were added to cooled mixtures of acetic acid (50%, 50 ml) and iodine (0.04 *M*) in potassium iodide solution (0.1 *M*, 50 ml), and excess iodine was titrated against standardized sodium thiosulphate solution (0.05 *N*) after 5 minutes.

3. The gases evolved from similar aliquots (using 150 ml of 50% phosphoric acid) were removed by a stream of nitrogen and passed through a solution of cadmium chloride (0.1 *M*) and sodium acetate (0.1 *M*) at pH = 6.8 to absorb hydrogen sulphide and subsequently through methanolic potassium hydroxide to absorb carbon disulphide (13). The sulphur was estimated either by the barium sulphate method (method 1) or by the iodometric method (method 2). Carbonyl sulphide was found to precipitate cadmium sulphide from solutions buffered at pH 4–7 with acetic acid and sodium acetate (26), and its presence invalidates the method.

4. The cellulose in aliquots (40 ml) was estimated by oxidation (27) with potassium dichromate (1 *N*) and sulphuric acid (70% in reacting solutions). Excess dichromate was titrated against ferrous ammonium sulphate using platinum and calomel electrodes.

Measurement of Ultraviolet Spectra

Aliquots were accurately diluted in sodium hydroxide solution (8%) to give a suitable range of concentrations, and the extinction for maxima at about 380 and 300 mμ was measured using 1-cm silica cells in a Beckman D.K. 2 spectrophotometer. The calculated extinction coefficient evidently depended upon the slit width which adjusts automatically at a given sensitivity.

Potassium Cellulose Xanthate

Samples of dissolving pulp (25 g) were steeped in potassium hydroxide solution (40%) at 5° C for 3 hours, filtered, and subjected to an air current for 3 hours. The alkali cellulose (35 g) and potassium hydroxide solution (100–400 ml, 40%) were stirred with carbon disulphide (400 ml) for 16–40 hours. Excess carbon disulphide was removed under vacuum (15 mm), and samples were examined as described for viscose.

Sodium Cellulose Xanthate

Viscose (50 g; S, 1.45; cellulose, 6.78%; D.S., 0.51), diluted with ice water (250 ml), was titrated with cold acetic acid (6 *M*) to pH 6–7 using glass and calomel electrodes. Cold acetone (750 ml), run in over about 10 minutes, gave a precipitate which was either centrifuged or filtered off and ground successively in a mortar with cold acetone (3×50 ml), acetone–ether mixtures (40:10, 30:20, 20:30, 10:40), and finally with anhydrous cold ether (3×50 ml). Samples for analysis (methods 1 and 4) and spectroscopic examination were dried at 20° C (0.5 mm Hg over phosphorus pentoxide). The use of methanol in place of acetone gave a more gelatinous precipitate, which was more difficult to filter or centrifuge, and which resulted in extensive xanthation of the methanol (λ_{max} = 380, 300 mμ). Dissolution of these products in cold sodium hydroxide solution formed more trithiocarbonate (λ_{max} 500 mμ; shoulder 325 mμ). Sample A, Fig. 4, found: S, 9.27; cellulose, 62.6%; D.S., 0.38.

S-*Methyl Cellulose Xanthate*

S-Methyl cellulose xanthate (cell—O—CS—SCH_3) was prepared from the precipitated sodium cellulose xanthate sample (D.S., 0.38) and methyl iodide, essentially as recently described (17). The product was washed successively with acetone, water, acetone, and solvent exchanged into ether. Sulphur in samples dried to constant weight at 20° C (0.5 mm Hg) was determined by use of a Parr bomb (28). All samples were insoluble in pyridine and lost sulphur in this solvent at 100° C (29). Sample B, Fig. 6B, found: C, 39.05; H, 5.57; S, 9.01. Calculated for D.S. of 0.26: C, 42.19; H, 5.71; S, 9.01%.

Diethylacetamido Cellulose Xanthate

Diethylacetamido cellulose xanthate (cell—O—CS—S—CH_2—$CONEt_2$) was precipitated from the same sample (D.S., 0.51) of diluted viscose neutralized to pH 6.5 as described (13), and dried to constant weight at 20° C (0.1 mm Hg). Sulphur (28) and nitrogen (30) were estimated by standard methods. This substance lost sulphur in pyridine at 100° C. Sample C, Fig. 4, found: S, 11.94; N, 2.36%; D.S., 0.47; 0.40 respectively. After reprecipitation from dimethyl sulphoxide by ether, found: S, 10.2%; D.S. 0.40.

Cellulose Dixanthate

Cellulose dixanthate (cell—O—CS—S—)$_2$, obtained in iodometric titration (method 2) of the same viscose sample, was filtered, washed with solvents (as for cellulose xanthate), and dried at 80° C (0.5 mm Hg). Sulphur (28) determinations showed loss of xanthate. Found: C, 40.73; H, 5.81; S, 9.93. Calculated for D.S. of 0.29: C, 40.64; H, 5.33; S, 9.93%.

Potassium Ethyl Xanthate

Potassium ethyl xanthate was recrystallized from acetone and dried at 80° C (0.1 mm Hg). Sulphur, determined by methods 1, 2, and 3, gave 39.3, 39.13, and 39.01% respectively (purity 98%).

Potassium Ethyl Monothiocarbonate

Potassium hydroxide (6.6 g) in pure ethanol (30 ml) at −15° C absorbed less (5½ liters) than the calculated amount of pure carbonyl sulphide (11 liters) and precipitated a white solid A (10 g), which was filtered off and rapidly extracted with ethanol (50 ml) at 50° C, leaving a hygroscopic powder B (5.3 g, S 2.39%) containing much carbonate. The ethanol extract, cooled to 5° C, gave white needles C (2.6 g, S 18.54%), which on similar extraction left an insoluble powder (10 mg); cooling to 0° C gave white needles D (1 g). Found S: 21.34% C_2H_5COS; K requires S, 22.22%, purity 96%. The infrared spectra (Fig. 4) of these samples showed bicarbonate as an impurity which was not completely removed. The mother liquor showed absorption maxima at 380 and 300 mμ, which are characteristic of the ethyl xanthate.

Four similar preparations and the use of sodium hydroxide, sodium ethoxide in toluene, and potassium hydrosulphide with ethylchloroformate in ethanol gave similar mixtures. Extraction with ethanol at lower temperatures (−5 to −10° C) gave small amounts of impure products. Previous accounts of this preparation (e.g. 31) apparently resulted in a purer product. Samples of our crude reaction product A, analyzed by methods 1, 2, and 4, evidently contained sulphide as indicated by the higher iodometric results. The fact that carbonyl sulphide reacted with the cadmium reagent invalidated the gas analytical method (sulphur in sample A: by Ba_2SO_4, 10.87%; by iodometry, 11.9%; as COS^-, 9.91%; as $S^=$, 0.96%; from gas analysis: absorbed by $Cd(OAc)_2$, 1.44%; absorbed by KOH/CH_3OH = 9.12%).

Sodium Cellulose Monothiocarbonate

Alkali cellulose from cotton linters and sodium hydroxide solution was treated with liquid carbonyl sulphide at −70° C and the temperature was allowed to rise to −5° C over 18 hours. Weighed samples of the white solid were analyzed and partial solutions in 8% purified sodium hydroxide were prepared as described for viscose. Solutions in aqueous sodium hydroxide showed no absorption maxima at 382 and 302 mμ, indicating the absence of xanthate. Samples prepared by passing carbonyl sulphide gas over alkali cellulose at −15° C were green in color only when carbon disulphide was present in the carbonyl sulphide. The green color was extracted into acetone at −15° C under nitrogen, and the spectrum (Fig. 3) was examined in 1-cm silica cells; in air the color faded, and sulphur precipitated.

Diethylacetamido Cellulose Monothiocarbonate, Attempted Preparation

Sodium cellulose monothiocarbonate (10 g) in cold sodium hydroxide solution (100 ml) and N,N-diethylchloroacetamide (5 ml) were mixed and the solution was neutralized to pH 7 with cold acetic acid (6 *N*) using glass and calomel electrodes. After 2 hours, the gelatinous precipitate (3.9 g) was filtered, washed with water and solvents (13), and dried at 28° C over phosphorus pentoxide (0.1 mm Hg). The product contained neither sulphur nor nitrogen. Found: C, 40.25, 42.43; H, 6.63, 6.41. Calculated for $(C_6H_{10}O_5)_n$: C, 44.4; H, 6.22%.

Infrared Examination

Solid samples (1.5 mg) were dried at 20° C under 0.1 mm Hg pressure over phosphorus pentoxide, finely divided and made into pellets with potassium bromide (300 mg) (33). Freshly prepared cellulose monothiocarbonate samples containing appreciable quantities of water were examined as pellets and as Nujol mulls. All spectra were measured with a Perkin–Elmer model 13 infrared spectrophotometer equipped with NaCl optics and operated in double-beam mode.

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THE IDENTITY OF BAYERITE-*a* AND BAYERITE-*b*

G. YAMAGUCHI AND K. SAKAMOTO

In 1928 Fricke (1) reported the existence of two types of bayerite, $Al_2O_3 . 3H_2O$, which he called *a* and *b*. Both give almost identical X-ray diffraction powder patterns except that the relative intensities of some of the lines are not the same in both patterns. The most striking example is the 00.1 line, which is one of the strongest in the pattern of type *b* and is a little stronger than the 10.0 line, while in type *a* it is conspicuously weaker than the latter. Nishimura (2) has recently obtained the less common type *a* while preparing Raney nickel from a Ni–Al alloy. He has confirmed the very weak intensity of the 00.1 line and has found that it varies in relative intensity according to the conditions under which the bayerite is formed. We have repeated the preparation of type *a* by Nishimura's method and have obtained type *b* by a procedure reported elsewhere (3). Both types have now been examined by the X-ray diffraction powder method (Fig. 1), by electron microscopy (Fig. 2), and by electron microdiffraction techniques (Fig. 3). The observed *d* spacings, derived from the diffractometer charts reproduced in Fig. 1, are listed in Table I,

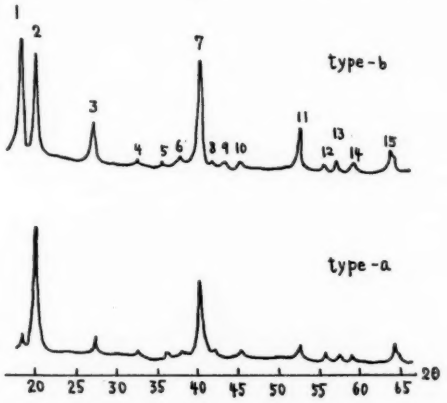

FIG. 1. X-Ray diffraction patterns of bayerite-*a* and bayerite-*b* (Cu $k\alpha_1$, filtered; Norelco-type diffractometer).

where it will be seen that, except for relative intensities, the powder patterns of type *a* and type *b* are identical. The electron micrographs (Fig. 2) show that type *a* crystallizes with a needle-like, or lath-like, habit and a tendency towards parallel growth along the direction of the needle axis, while type *b* crystallizes in thin, hexagonal plates. The structural periodicity shown in the electron microdiffraction photograph of a single crystal of type *a* (Fig. 3) was found to be 4.71 Å by comparison with an electron diffraction photograph of gold foil; this value is in good agreement with the length of the *c* axis of type *b* (a = 5.047 Å, c = 4.730 Å; S.G., $P\bar{3}1m$) reported elsewhere (4). These results indicate that bayerite-*a* and bayerite-*b* are structurally identical, that the crystals of type *a* are elongated in the direction of the *c* axis, and that the observed differences between types *a*

TABLE I

X-Ray diffraction powder data for bayerite-*a* and bayerite-*b*

	Bayerite-*a*		Bayerite-*b*			
Line	d_{obs}	I/I_1	d_{obs}	I/I_1	d_{calc}*	*hk.l*
1	4.745	7	4.745	100	4.730	00.1
2	4.381	100	4.385	90	4.371	10.0
3	3.209	9	3.210	34	3.210	10.1
4	2.706	2	2.705	2	—	—
—	—	—	—	—	2.523	11.0
5	2.468	4	2.468	2	—	—
6	2.365	1	2.368	7	2.365	00.2
7	2.222	62	2.227	90	2.226	11.1
—	—	—	—	—	2.185	20.0
8	2.176	3	2.171	2	—	—
9	—	—	2.079	3	2.080	10.2
10	1.989	4	1.987	4	1.984	20.1
11	1.722	11	1.725	45	1.726	11.2
12	1.647	4	1.653	3	1.652	12.0
13	1.605	4	1.604	10	1.605	20.2
—	—	—	—	—	1.577	00.3
14	1.558	5	1.560	15	1.560	12.1
—	—	—	—	—	1.483	10.3
15	1.458	6	1.460	25	1.457	30.0

*d_{calc} from $a = 5.047$ Å, $c = 4.730$ Å for bayerite-*b* (4).

and *b* in the relative intensities of some of the X-ray diffraction powder lines, particularly the abnormally weak 00.1 line of type *a*, are the result of preferred orientation of the crystals of type *a* parallel to their length.

Grateful acknowledgment is made to Dr. S. Nishimura for helpful advice regarding the preparation of bayerite-*a*.

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DIPHENYLCYANAMIDE DERIVATIVES. IV. 1,1-bis(4-CHLOROPHENYL)-UREA*

J. R. Robinson

The synthesis of 1,1-bis(4-chlorophenyl)-urea was reported as resulting from the aqueous alkaline hydrolysis of 4,4′-dichlorodiphenylcyanamide; it was described as a white solid, m.p. 220°–222° (1). Subsequent investigation has shown this material to be a complex mixture containing 1,1-bis(4-chlorophenyl)-urea, some 4,4′-dichlorodiphenylamine, and a polymeric compound containing silicon leached from the Pyrex flask. (No

*Issued as contribution No. 169.

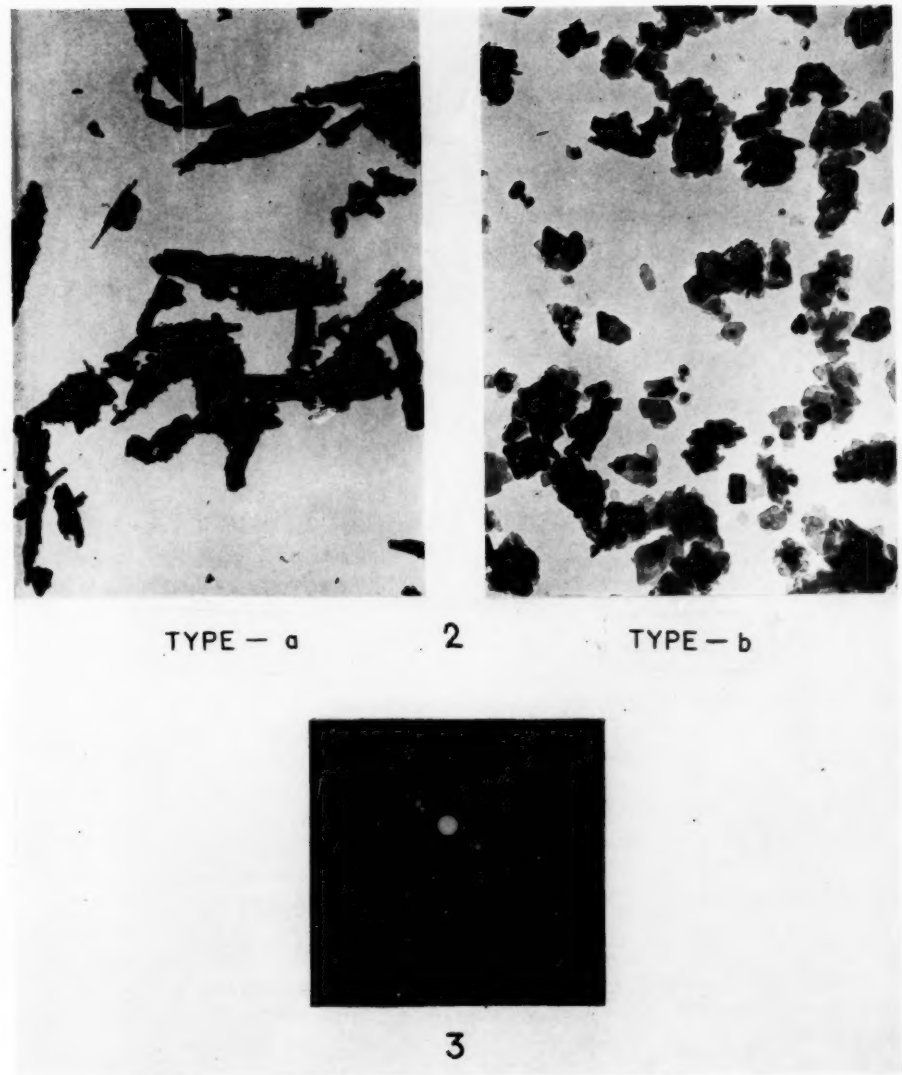

FIG. 2 Electron micrographs of bayerite-*a* and bayerite-*b* (×27,000).
FIG. 3. Electon microdiffraction pattern of bayerite-*a* (acceleration voltage 100 kv).

silicone grease was used.) The purpose of this note is to record a corrected description of 1,1-bis(4-chlorophenyl)-urea in the pure state.

This unsymmetrically substituted urea has now been obtained by a different method and characterized through conversion to known derivatives by the well-established reactions illustrated in Fig. 1.

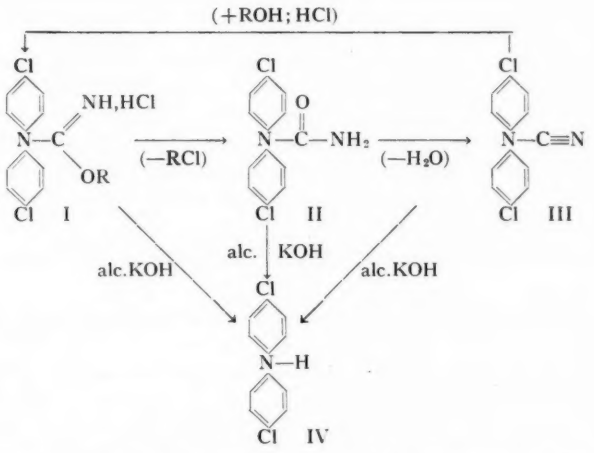

FIG. 1. The synthesis and hydrolysis of 1,1-bis(4-chlorophenyl)-urea.

Although the 3,3-diphenyl-2-alkylpseudoureas are known to undergo rearrangement and cleavage on heating to form diphenylamine and the corresponding alkyl isocyanate (2), their hydrochloride salts, when heated, retain the urea skeleton resulting in the formation of unsymmetrical diphenylurea and the corresponding alkyl chloride (3). Similar behavior of the 3,3-dialkyl-2-alkylpseudoureas was reported by McKee (4). This property has provided a convenient method for synthesis of 1,1-bis(4-chlorophenyl)-urea (II, Fig. 1). The reaction of water with the pseudourea salts, at room temperature, is observed to produce the same material, though slowly.

Evidence for the structure of II is provided by its ready dehydration to 4,4′-dichlorodiphenylcyanamide (III) using Kurzer's technique (5), as well as by its hydrolysis to 4,4′-dichlorodiphenylamine (IV) in alcoholic potassium hydroxide. Further, the infrared spectrum* of the material in the region 3600–3200 cm^{-1} is identical with that of 1,1-diphenylurea and 1,1-bis(4-methoxyphenyl)-urea, each having absorption peaks characteristic of $—NH_2$ at 3540 and 3430 cm^{-1}, as well as a strong peak at 1675 cm^{-1} typical of an amide carbonyl; a peak at 830 cm^{-1} is further indication of the paradisubstituted phenyl groups in this material.

The degradation of the 3,3-bis(4-chlorophenyl)-2-alkylpseudourea salts (I) to 4,4′-dichlorodiphenylamine may also be accomplished directly in boiling alcoholic potassium hydroxide; 4,4′-dichlorodiphenylcyanamide undergoes similar conversion (1).

Acknowledgment is made to Mr. R. W. White of this laboratory for recording and interpreting the infrared spectra of these materials.

EXPERIMENTAL*

3,3'-Bis(4-chlorophenyl)-2-alkylpseudoureas

3,3-Bis(4-chlorophenyl)-2-alkylpseudoureas (Fig. 1, I where R = C_1 to C_{12} alkyl groups) were prepared by the alkoxide catalyzed addition of alcohols to 4,4'-dichlorodiphenylcyanamide (1, 3).

1,1-Bis(4-chlorophenyl)-urea

(a) By Thermal Decomposition of 3,3-Bis(4-chlorophenyl)-2-n-heptylpseudourea Hydrochloride

The pseudourea salt (2.0 g, 4.8 mmoles) was spread thinly over a watch glass and kept in a 90° oven until odorless and no longer water-soluble (5 days). The residue was recrystallized from methanol–water, ethanol–water, then dried *in vacuo* over calcium chloride. The yield of purified material (fine, colorless needles) was 1.0 g (75%); it gave a positive test for organic halogen, a negative test for halide, and melted at 187°–188° C. Its infrared absorption characteristics are described above. Anal. Calc. for $C_{13}H_{10}N_2OCl_2$: C, 55.53; H, 3.58%. Found: C, 55.75; H, 3.57%.

(b) By Thermal Decomposition of 3,3-Bis(4-chlorophenyl)-2-n-butylpseudourea Hydrochloride

The salt (1.17 g, 3.18 mmoles) was kept at the melting point, under vacuum (0.2 mm of mercury), on a manifold to which a cold trap was attached. After 90 minutes the crude residue weighed 0.87 g (97%) and repeated recrystallization from ethanol–water gave fine, colorless needles, m.p. 187°–188°.

The distillate in the cold trap weighed 0.265 g (90%) and had the boiling point (77°–78°) and refractive index (1.4014) of *n*-butylchloride.

(c) By the Action of Water on 3,3-Bis(4-chlorophenyl)-2-methylpseudourea Hydrochloride

An aqueous solution of the salt (0.5 g, 1.5 mmoles in 100 ml of water) was left standing in the dark at room temperature. After several days, long needle-like crystals began to grow. After 7 months the material was filtered off and recrystallized from methanol–water and ethanol–water, giving colorless needles, m.p. 187°–188°.

The three products (*a*, *b*, and *c*) were identical in chemical and physical properties.

4,4'-Dichlorodiphenylcyanamide

A solution of 0.84 g (2.98 mmoles) of 1,1-bis(4-chlorophenyl)-urea (above) in 25 ml of pyridine was treated with 1.77 g (10 mmoles) of benzenesulphonyl chloride and left standing at room temperature for 12 hours. The reaction product, after precipitation with water, filtration, and washing, was recrystallized from methanol yielding 0.58 g (73%) of long colorless needles of 4,4'-dichlorodiphenylcyanamide, m.p. 114°–115°. The material was confirmed by a mixed melting point with an authentic sample (1).

4,4'-Dichlorodiphenylamine

A 5-ml methanolic potassium hydroxide solution of 1,1-bis(4-chlorophenyl)-urea (0.12 g, 0.42 mmole) was boiled under reflux for 24 hours; some ammonia was evolved. After crystals of potassium carbonate were filtered off, the product was precipitated by adding water and recrystallized from 110° petroleum ether yielding 0.085 g (85%) of 4,4'-dichlorodiphenylamine, colorless needles, m.p. 77°–78°.

The identical product was obtained in 80–90% yield, by the same technique, starting with any of the 3,3-bis(4-chlorophenyl)-2-alkylpseudourea hydrochlorides.

**Melting points were obtained using a Fisher–Johns block and are uncorrected. Carbon and hydrogen analyses were by J. F. Alicino, Metuchen, N.J.*

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